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CHEMISTRY 30

MODULE 3:
ELECTROCHEMISTRY - GIVE AND TAKE

Distance Learning

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Chemistry 30

Module 3

Electrochemistry – Give and Take

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Alberta
EDUCATION

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Chemistry 30
Student Module
Module 3
Electrochemistry – Give and Take
Alberta Distance Learning Centre
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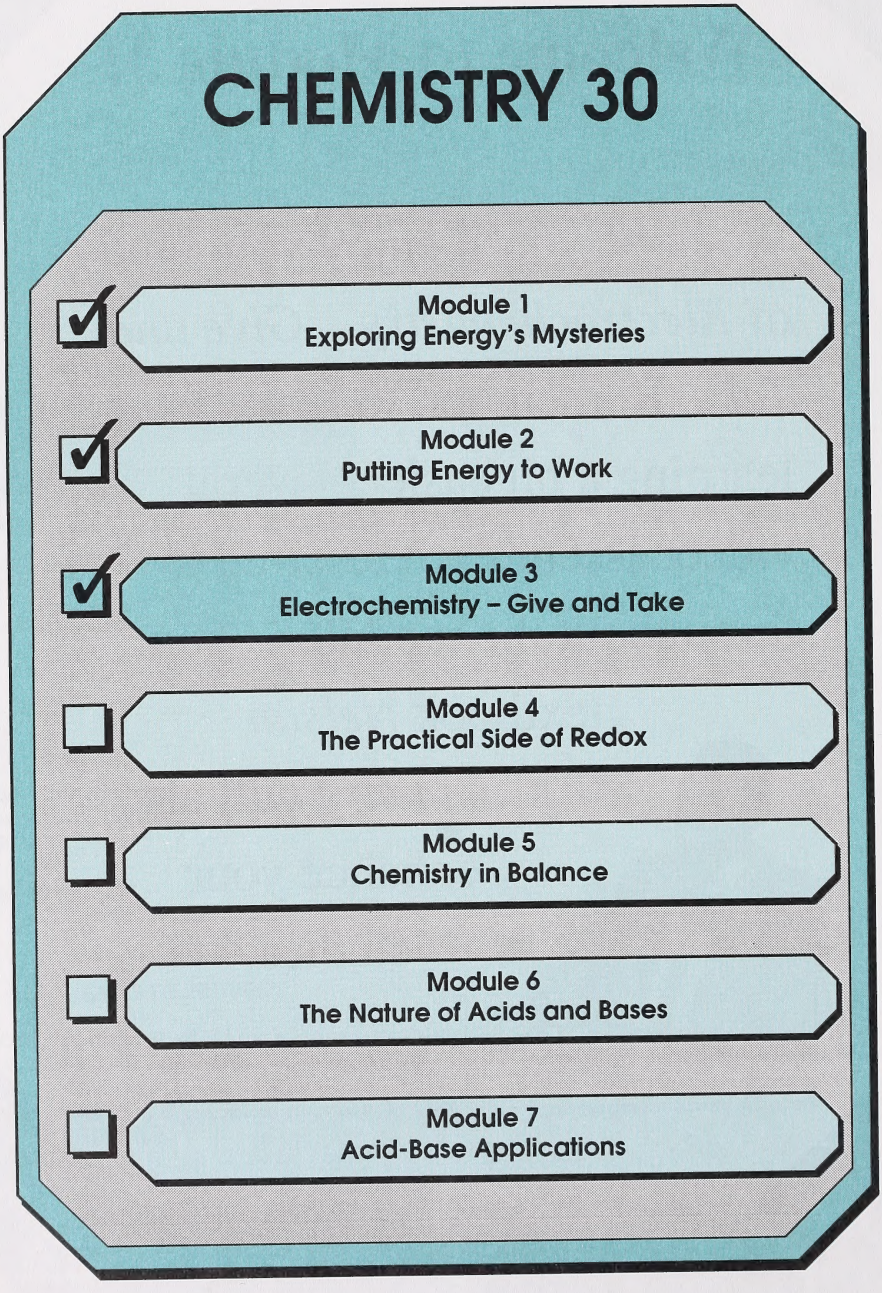
Welcome to Module 3!

We hope you'll enjoy your study of *Electrochemistry – Give and Take*. Remember, to make your learning a bit easier, several videocassettes are referred to throughout the modules. Your textbook *Nelson Chemistry* will also enhance your studies.

COURSE OVERVIEW

This course contains seven modules. The module you are working in is highlighted in deeper colour. It is recommended that you work through the modules in the order given since several concepts build on each other as you progress in the course.



MODULE 3

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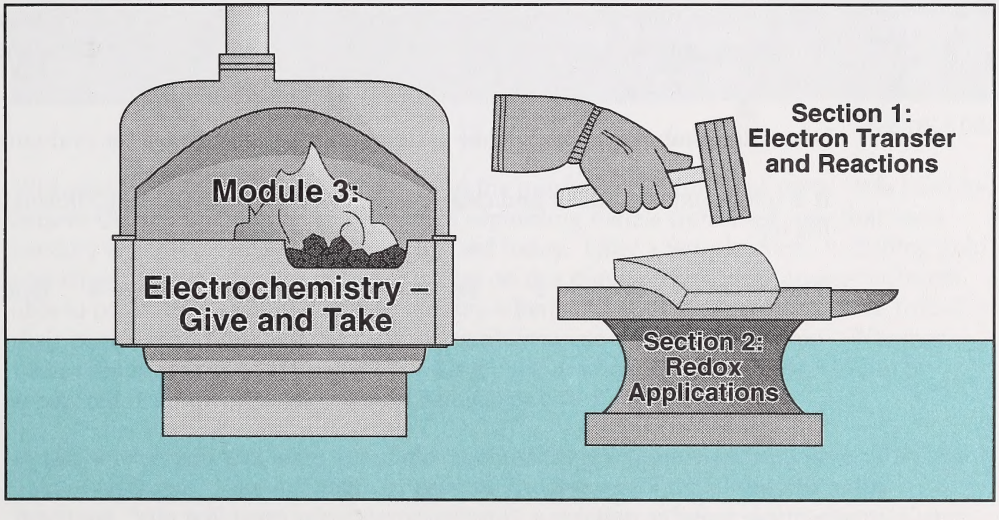
MODULE OVERVIEW

Sitting on his throne, overlooking the battle in the valley below, the ancient king knew his time was coming to an end. His mighty army, once feared by all in this part of the world, was being shredded by the newcomers. It was not a matter of numbers or strength of men, but a matter of passing technology.

Just as the king's ancestors had discovered a new way of extracting and forging a stronger metal for their weapons, it was now time for another passing. The invaders had a technology superior to his people's. With the invaders' metallurgical advances, through even stronger and lighter alloys, the king's doom was inevitable.

Throughout the history of humankind, civilizations have come and gone. Society's progression through thousands of years has been based on the ability to capture the wonders of science. The electrochemical process of metallurgy was responsible for this king's conquests as well as his downfall.

In this module you will be learning about the building blocks of electrochemical reactions. These types of reactions are not only involved in metallurgy, but they also form the basis for processes such as photosynthesis, combustion, respiration, and metabolism. You will apply the basic concepts of electrochemical reactions to these processes and further develop your knowledge of electrochemistry in preparation for the applications presented in Module 4.



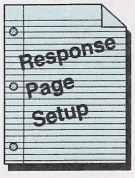
Evaluation

Your mark in this module will be determined by how well you complete the assignments at the end of each section. You must complete all assignments. In this module you are expected to complete three assignments. The mark distribution is as follows:

Section 1 Assignment	**40 marks**
Section 2 Assignment	**40 marks**
Final Module Assignment	**20 marks**
TOTAL	**100 marks**

When doing your assignments, work slowly and carefully. If you are having difficulties, go back and review the appropriate section.

Read all parts of your assignment carefully. Plan and do your rough work on your own paper. Revise and edit your responses; then set up your final copy for submission on your own paper. Lined looseleaf is recommended. Make sure your answers are neat and organized, with wide left margins and space for teacher comments after each assignment.



When you see this icon, ideas and details are provided to help you set up and organize your answer in a certain way.

Before submitting your responses, be sure to proofread them carefully to ensure they say what you want, that they are neat and clear, and that they are complete and missing no material.

You will be submitting **only** your **assignment response pages** for evaluation.

It is important to number and clearly identify each page with this information placed at the top.

Chemistry 30 – Module 3 Section # Assignment Page # Name and ID #

SECTION 1

Electron Transfer and Reactions

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With society's modern-day knowledge in the mining and refining of metals it is hard to believe that the same basic techniques of separating metals from their ores that were used by ancient cultures are still being used today. Only a few elements, including gold and silver, can be found in their pure form on this planet. How were ancient cultures able to produce copper and iron weaponry when neither of these metals can be found in their pure form? How are the various metals you use each day produced? Whether these metals are used in jewellery, cooking pots, or in construction, they have to be separated from their ores by electrochemical or redox reactions.

In this section you will learn about electrochemical reactions, which involve a loss and gain of electrons. You will begin by relating the process of metallurgy to redox reactions. You will learn what factors identify a reaction as being electrochemical and develop the methods required to write and balance redox equations. Finally you will examine the process used to predict which redox reactions will most likely occur in a given situation.

ACTIVITY 1

In the Beginning

1. How do you suppose the first metals were discovered on Earth?

Check your answers by turning to the Appendix, Section 1: Activity 1.

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Have you ever wondered about where many of the metals you come in contact with every day come from? What about the gold in the ring you wear, or the aluminum in the cooking pot you may have used recently? If you think about it, there are many things around you that are made from metals.

The ancient peoples of the world learned very quickly how to isolate metals from their ores. They had no scientific background for the isolating of metals. The fact that electrons were exchanged when a redox reaction took place was of no importance to these people. They were able to produce copper, tin, and iron by heating the ores of each of these metals in the simple furnaces they made. They progressed to combining molten metals to produce alloys like bronze for their everyday needs.

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In this activity you will learn about the basic concepts behind metallurgy. To learn more about how metals were produced in the past and how those methods link to the present, read pages 346-348 of your textbook.

reduction – the process by which electrons are gained in a reaction

oxidation – the process by which electrons are lost in a reaction

You learned about the terms *reduction* and *oxidation* in Chemistry 20. You will now revisit these terms and see how they will become important to you in this chemistry module. Follow up on the reading by answering the following questions.

2. If the king described in the Module Overview was from the bronze age, what metal was used in the weapons of the invading army?

3. a. The invaders' advanced metallurgical technology allowed them to take large amounts of iron ore and produce smaller amounts of iron by a process called _____.

   b. Name two other common metals that were produced by this process.

4. To produce the metals that were used to make weapons, a substance had to be used to help the reaction occur. This substance would generally be called a _____.

5. Name two chemical substances other than oxygen that can cause a metal to be oxidized.

Check your answers by turning to the Appendix, Section 1: Activity 1.

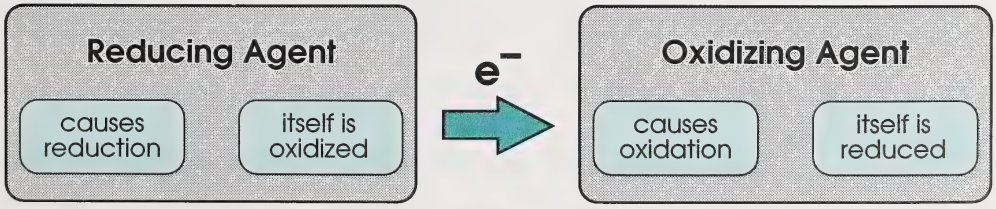
oxidizing agent – a chemical species that accepts electrons in a reaction

reducing agent – a chemical species that donates electrons in a reaction

In the reading you reviewed the terms *reduction* and *oxidation*. You were also exposed to two new terms – *oxidizing agent* and *reducing agent*. To help you understand the difference in the terms, think of yourself planning a vacation. You wish to travel to a tropical island, but you need help organizing your trip. You would acquire the services of a travel agent. With the help of the agent you make the trip.

Similarly, a metal ore that needs to be reduced requires a reducing agent to complete the reaction. When a chemical species is to be reduced it will gain electrons, and it needs the reducing agent to supply these electrons.

Just as *reduction* and *reducing agents* are terms that are important to the chemistry of an electrochemical reaction, so too, are the terms *oxidation* and *oxidizing agents*. A chemical species that undergoes oxidation is losing electrons, while the oxidizing agent is the species that is accepting the electrons. At one time any reaction that involved oxygen came to be called an oxidation reaction. You have already seen, though, that other substances can act as oxidizing agents in a reaction.



For a chemical change to be a redox reaction, you must always have both an oxidizing agent and a reducing agent. That also means that both oxidation and reduction must occur.

You have probably seen a good example of a common redox reaction that occurs most everywhere. The oxidation of iron, called corrosion or rusting, is a process that causes millions of dollars in damage each year. The iron loses electrons to substances such as water and oxygen in the air. These substances, in turn, accept electrons from the iron, thereby causing iron to form a compound called rust.

To become more familiar with the terms and process just discussed, do the following question.

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Nelson CHEMISTRY

6. Do textbook question 2 (a, b, c, d, e, and f) on page 348 of *Nelson Chemistry*.

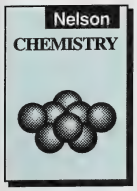
Check your answers by turning to the Appendix, Section 1: Activity 1.

One of the best ways you can confirm your understanding of electrochemical processes is to perform a lab investigation related to redox concepts. The following investigation not only reviews chemical equations you have written in the past, but demonstrates how your redox knowledge ties in with these equations.

Important Note – Disposal Information

Any chemicals used in investigations in this module should not be poured down the sink. Keep waste materials in closed, labelled containers. Contact your nearest high school or fire department for disposal information.

Investigation 12.1: Single Replacement Reactions



Carefully follow the directions given for this investigation on page 349 of your textbook. Pay special attention to the required components, safety aspects, and applied science skills.

Science Skills

- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Guidelines

- Use 0.10 mol/L solutions.
- Use pieces of metal approximately 1 cm in length.
- Use a piece of steel wool to clean the surface of the metal strips.
- The test tubes need only be one-third filled with each solution used.

7. Predict a complete balanced equation, including states, for each of the single replacement reactions given in the problem of this investigation.

8. Write a brief description of your design, keeping in mind the directions given in your textbook in the introduction of this investigation.

PATHWAYS

If you have access to supervised laboratory facilities, do Part A.
If you do not have access to laboratory facilities, do Part B.

Part A

9. Prepare an evidence table similar to the following, perform the procedure, and record your observations.

Test Tube Number	Initial Observations after Reaction	Observations from Diagnostic Tests
1		

Check your answers by turning to the Appendix, Section 1: Activity 1.

You will complete your analysis by doing the questions after Part B.

End of Part A

Part B

You will not be able to perform the following parts of the investigation:

- the second and fifth reactions as indicated in the problem of this investigation
- the diagnostic tests for halogens

Directions for how you can proceed are given in the following question.

10. Prepare an evidence table similar to the following, including all five test tubes. Copy the given data for Reactions 2 and 5. Perform the remaining parts of the investigation and record your observations.

Test Tube Number	Initial Observations after Reaction	Observations from Diagnostic Tests
1		
2	There was a yellow-brown coloured solution.	The solvent layer was orange in colour.
3		
4		
5	There was a yellow-brown coloured solution.	The solvent layer was violet (purple) in colour.

Check your answers by turning to the Appendix, Section 1: Activity 1.

Complete your investigation by doing the following analysis questions.

End of Part B

Analysis

11. Complete the following table using the evidence you gathered.

Test Tube Number	Analysis – Confirmation of Products
1	

12. Evaluate the design and procedure of your investigation based on your results.

Check your answers by turning to the Appendix, Section 1: Activity 1.

In Chemistry 20 you did a brief study of oxidation, reduction, and the writing of **half-reaction equations**. You may wish to review these concepts on page 211 of your textbook before doing the next question.

half-reaction equations – balanced equations which describe one-half of an overall process

13. To complete your analysis for Investigation 12.1, answer textbook questions 6, 7, and 8 on pages 349-350 of *Nelson Chemistry*.

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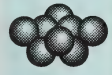

Check your answers by turning to the Appendix, Section 1: Activity 1.

Single replacement reactions are examples of electrochemical reactions in which a loss and gain of electrons takes place. In all of the reactions you performed in the preceding investigation, the products formed were due to a loss or a gain of electrons.

Example

$Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$

lost electrons ↓ $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^-$ ↓ electrons on right

gained electrons ↓ $Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$ ↑ electrons on left

OIL RIG
Oxidation **I**s **L**oss of electrons.
Reduction **I**s **G**ain of electrons.

Now that you have worked with half-reactions, you will move on to ideas that chemists use to explain how these half-reactions can combine to give an overall redox reaction.

ACTIVITY 2

Redox – What Is It?

The more practical (empirical) side of reduction and oxidation was demonstrated to you in Activity 1. You learned how ancient societies were able to rely on knowledge gained through observation to isolate metals for their use. But what about the isolation of the other elements? How were people able to develop the technology to get the large amounts of hydrogen gas needed for early air balloons like the zeppelins, for example?

In this activity you will learn about the theoretical side of what actually takes place when metals or nonmetals are formed from their naturally occurring compounds.

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Read pages 350-352 of your textbook to help you understand the theoretical ideas behind reduction and oxidation and how the terms are used to explain the loss and gain of electrons in redox reactions. Your textbook discusses the reduction of aqueous silver nitrate to silver metal in the presence of solid copper. Consider this example when answering the following questions.

1. In terms of electrons lost and gained, what happens to the copper metal?

2. In terms of electrons lost and gained, what happens to the silver ions?

3. Which species supplies electrons for reduction to take place?

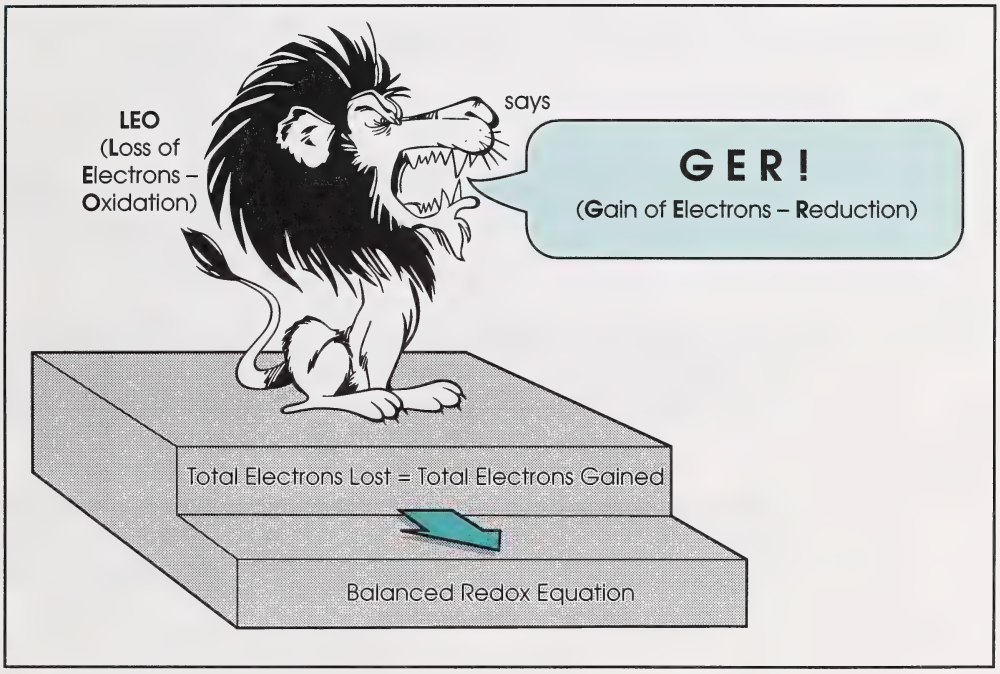
4. Which species gains the electrons so that oxidation can take place?

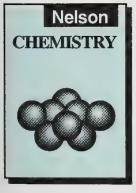
5. What is the relationship between the electrons lost and gained in the overall equation?

Check your answers by turning to the Appendix, Section 1: Activity 2.

mnemonic – a memory aid

As you discovered in Chemistry 20, you can make use of a **mnemonic** to help you remember what the terms *reduction* and *oxidation* mean – LEO the lion says GER, or simply, LEO says GER.

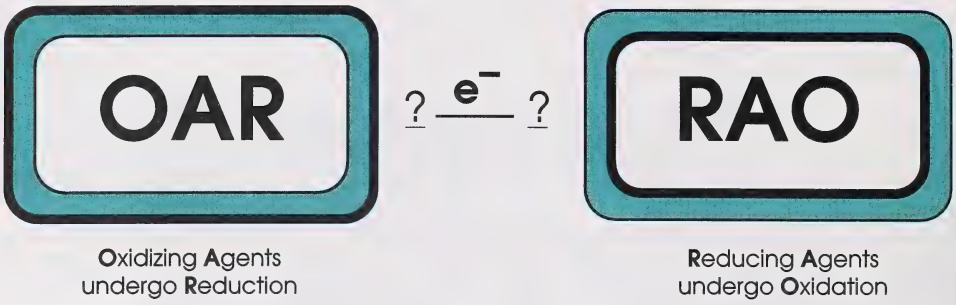




6. Referring again to the solid copper and aqueous silver nitrate reaction from your textbook (pages 350-351), what chemical species is associated with LEO? What species is associated with GER?

You learned in Activity 1 that reduction cannot occur without a reducing agent, just as oxidation cannot occur without an oxidizing agent. Like oxidation and reduction, the terms *oxidizing agent* (OA) and *reducing agent* (RA) can be easily confused with each other. To help keep them straight, remember the word OAR. It is another mnemonic – Oxidizing Agents undergo Reduction. The reverse of OAR is RAO, which is the mnemonic for Reducing Agents undergo Oxidation.

7. Reproduce the following graphic on your paper.



Place the missing arrowhead on the correct end of the line between the boxes to indicate the direction that electrons flow in a redox reaction.

8. Use the following single replacement reaction for this fill-in-the-blanks question.

$$Pb(NO_3)_{2(aq)} + Zn_{(s)} \rightarrow Pb_{(s)} + Zn(NO_3)_{2(aq)}$$

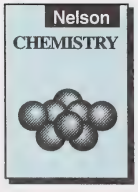
in solution ⇩ really looks like

$$Pb^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)} + Zn_{(s)} \rightarrow Pb_{(s)} + Zn^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)}$$

the net reaction (no spectators) ⇩ then becomes

$$Pb^{2+}_{(aq)} + Zn_{(s)} \rightarrow Pb_{(s)} + Zn^{2+}_{(aq)}$$

The zinc metal is being _____ (oxidized/reduced) by the oxidizing agent which is the _____ (lead ion/zinc metal). The zinc metal is undergoing a _____ (loss/gain) of electrons which is called _____ (oxidation/reduction). The lead ion is being _____ (oxidized/reduced) by the reducing agent which is the _____ (lead ion/zinc metal). The lead ion is undergoing a _____ (loss/gain) of electrons which is called _____ (oxidation/reduction).

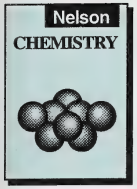


9. Do textbook questions 10 and 11 from pages 352-353 of *Nelson Chemistry*.

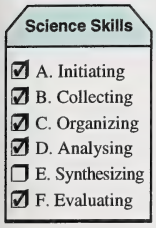
Check your answers by turning to the Appendix, Section 1: Activity 2.

You have just learned that for a transfer of electrons to happen there has to be both an oxidizing agent and a reducing agent. Does having an oxidizing agent and a reducing agent present mean that a redox reaction will automatically take place? You will discover the answer to this question by doing the following lab investigation.

Investigation 12.2: Spontaneity of Redox Reactions



Carefully follow the directions given for this investigation on page 353 of your textbook. Pay special attention to the required components, safety aspects, and applied science skills.



Guidelines

- Do not touch any of the chemical solutions. Silver nitrate, for example, will stain your skin black.
- Disposal Information: Any chemicals used should not be poured down the sink. Keep waste materials in closed, labelled containers. Contact your nearest high school or fire department for disposal information.
- All the solutions you will use are 0.10 mol/L.
- Use a spot plate in order to conserve solutions. One spot can be used for each reaction. Ten to fifteen drops of solution per spot should be sufficient.
- Use pieces of metal approximately 0.5 cm-1.0 cm in length.
- You will not be using silver metal as it is quite expensive. You can record that no reaction occurs when silver metal is involved.
- Note the distinction between zinc and lead as given in the materials list on page 353 of your textbook.
- Use steel wool to clean the surface of the metal strips.
- Signs of a reaction, if occurring, should become evident within two or three minutes.

10. Based on the introduction for this investigation as given in your textbook, make a prediction of the outcome.

11. Write a brief description of your design.

spontaneous – occurring naturally without outside help

12. Record your evidence in a data table as shown in Figure 3.1. Fill in the blank areas with the observations you note while performing the investigation. Write the letter *S* to stand for a **spontaneous** reaction (this means the reaction took place). Write the letters *NS* to stand for non-spontaneous (this means there was no evidence of a reaction). Your spot plate can be set up similar to your evidence table (minus the silver metal column).

EVIDENCE TABLE

Metals / Metal ions	$Cu_{(s)}$	$Pb_{(s)}$	$Zn_{(s)}$	$Ag_{(s)}$ (given data)
$Cu^{2+}_{(aq)}$				NS
$Pb^{2+}_{(aq)}$				NS
$Zn^{2+}_{(aq)}$				NS
$Ag^{+}_{(aq)}$				

FIGURE 3.1: Spontaneity of Redox Reactions

Check your answers by turning to the Appendix, Section 1: Activity 2.

13. a. Make a list of the metal ions, beginning with the one that reacted the most number of times and following through the list to the ion that reacted the least number of times.

    b. Write the half-reaction equation for each of the metal ions.

14. a. Make a similar list of the solid metal strips using their tendency to have reacted. Again, begin with the metal that reacted the most number of times and follow through to the one that reacted the least.

    b. Write the half-reaction equation for each of the metals.

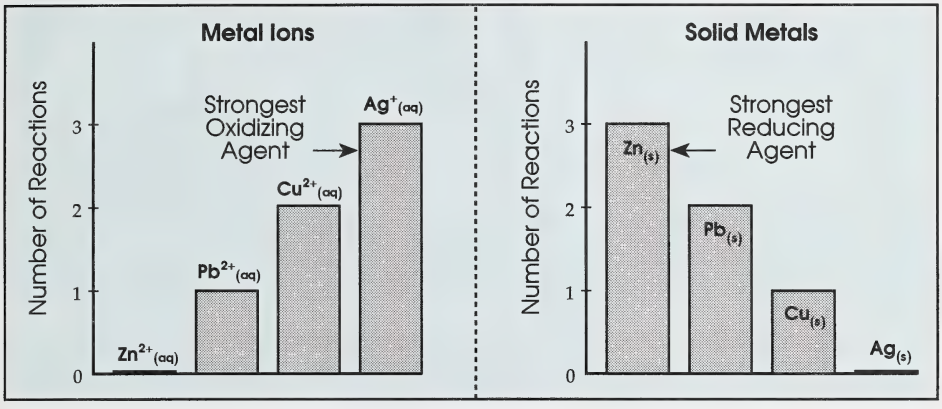
15. a. Considering your list of metal ion half-reaction equations, are the metal ions oxidizing agents or reducing agents?

    b. Were the metal ions themselves oxidized or reduced?

16. a. Considering your list of metal half-reaction equations, are the metals oxidizing agents or reducing agents?

    b. Were the metals oxidized or reduced?

INVESTIGATION 12.2: SPONTANEITY OF REDOX REACTIONS



Note that the metal ion which reacted the most number of times in this investigation is the strongest oxidizing agent. The solid metal that reacted the most number of times is called the strongest reducing agent.

17. You should now be able to answer the question posed before you started this investigation: Does having an oxidizing agent and a reducing agent present automatically mean that a redox reaction will take place?

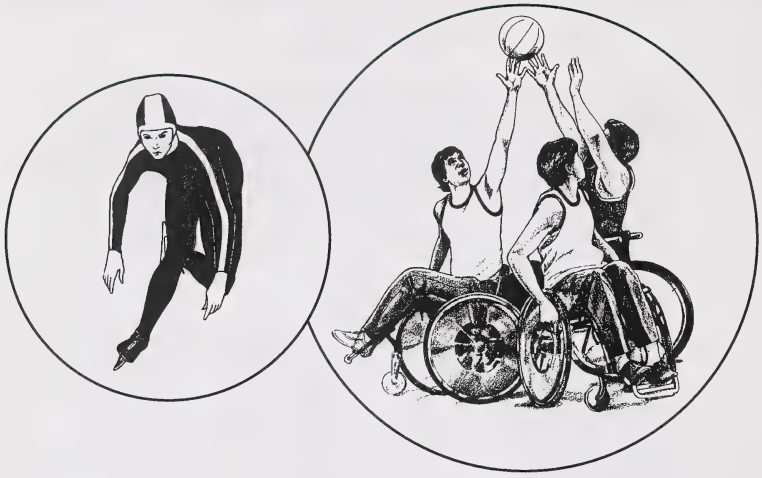
18. Write a brief evaluation of this investigation.

Check your answers by turning to the Appendix, Section 1: Activity 2.

In this activity you learned about the basic concepts that are important to the understanding of electrochemistry. These concepts and their associated terms – *reduction, oxidation, reducing agent, oxidizing agent,* and *spontaneity* – will, as you work through this module, help you to understand the mechanics of redox reactions.

ACTIVITY 3

Who's the Strongest?



When you see athletes participate in various activities, how do you know who will win? It is not an easy task to predict the winner. You may base your prediction on several criteria, one of which may be strength.

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You will find as you work through this activity that the winners in the competition for electrons by oxidizing and reducing agents are a little easier to predict. In Activity 2 you found a way experimentally to determine the strongest oxidizing agent and reducing agent. How do you decide which agents are the strongest, though, without doing an investigation? To help you decide, read page 354 in your textbook and answer the following questions.

1. Page 354 of your textbook discusses the results obtained from Investigation 12.2 (you performed the investigation in Activity 2).

   a. Based on the investigation results, which metal ion won the oxidizing agents' competition? How can you tell? What name is given to the winner?

   b. Based on the investigation results, which metal won the reducing agents' competition? How can you tell? What name is given to the winner?

Nelson CHEMISTRY

2. Answer textbook question 14 (a, b, c, d) on page 355 of *Nelson Chemistry*.

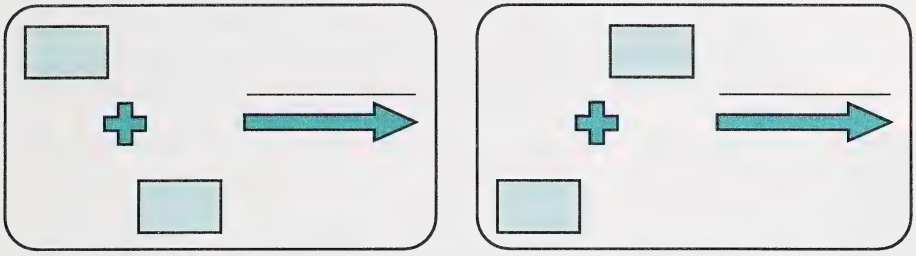
Check your answers by turning to the Appendix, Section 1: Activity 3.

Nelson CHEMISTRY

You have seen how regularities in evidence can be developed into a workable hypothesis. To see how the empirical hypothesis you developed can be turned into a generalization, read the middle section of page 355 in your textbook.

3. Briefly describe what your textbook calls the *redox spontaneity rule*.

4. Reproduce the following diagram and fill in the missing information from what you have read.



Check your answers by turning to the Appendix, Section 1: Activity 3.

Generalizations and hypotheses are useful to you only if they can be applied to new situations. To verify whether the rule for predicting spontaneity can apply to other redox reactions, do the following lab exercise.

Nelson CHEMISTRY

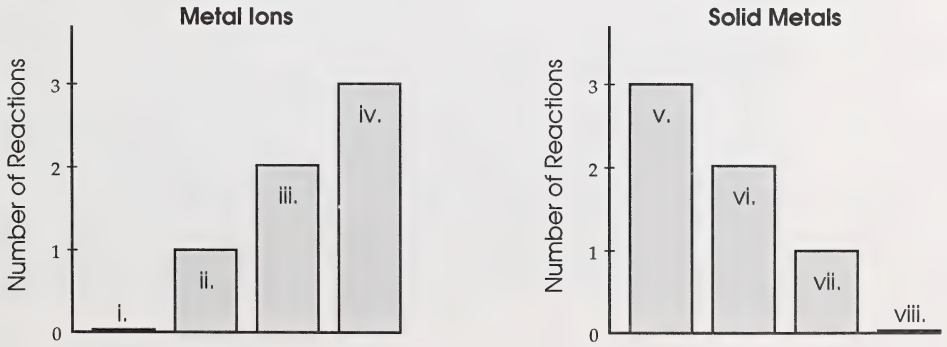
Lab Exercise 12A: Spontaneity of Reactions

Read through Lab Exercise 12A on pages 355-356 of your textbook and do the following three questions to complete your analysis and evaluation.

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

5. The following diagram illustrates the data given in the lab exercise. Identify the appropriate metals and metal ions according to the given information. Which species is the strongest oxidizing agent? Which species is the strongest reducing agent?



6. Develop a redox table as indicated in the lab exercise.

7. Write a brief evaluation, as directed in the lab exercise, of the spontaneity rule.

Check your answers by turning to the Appendix, Section 1: Activity 3.

There are several ways in which information about redox reactions can be given to you in order for you to construct a redox table. You have already created a redox table based on the number of reactions observed for each agent. Now you will see how a redox table can be developed from information given about the spontaneity of reactions tested.

Nelson CHEMISTRY

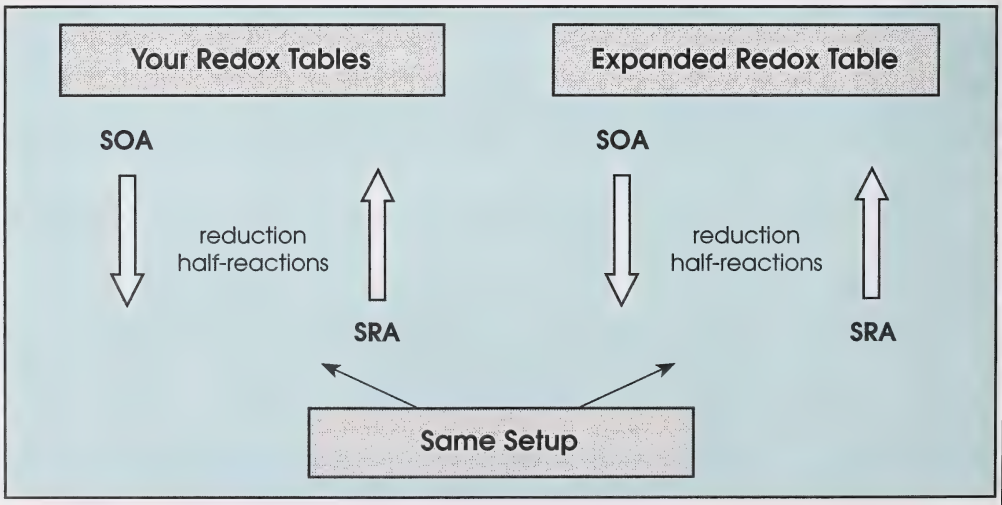
Read the bottom of page 356 and page 357 in your textbook and carefully study the given example.

8. To test your understanding of this alternate method of presenting information, do textbook questions 15, 16, and 17 on page 357 of your textbook.

Check your answers by turning to the Appendix, Section 1: Activity 3.

Nelson CHEMISTRY

Scientists have collected the data from many reactions similar to the ones you have studied and have looked at the relationships in strength between the oxidizing and reducing agents. The result of their work is an expanded redox table or reduction half-reactions table (see Appendix F on page 552 of your textbook).



As you can see, you will be able to use the expanded half-reactions table from your textbook to predict the relative strengths of many oxidizing and reducing agents. At times, you may have to use the redox table just to decide whether a substance is an oxidizing agent or a reducing agent. A substance found on the left side of the table is an oxidizing agent. A substance found on the right side of the table is a reducing agent. Remember to look at **both** sides of the table – many substances are found on both sides. These substances can be oxidizing agents **and** reducing agents.

Important Correction

Nelson CHEMISTRY

The *Redox Half-Reactions* table on page 552 of your textbook has an error that needs correcting. The dichromate ion half-reaction (eighth from the top) should read as follows:

$Cr_2O_7^{2-}(aq) + 14\ H^+(aq) + 6\ e^- \rightleftarrows 2\ Cr^{3+}(aq) + 7\ H_2O_{(l)}$ **+1.33**

↑ **correction**

The correction means that this (the dichromate) half-reaction should switch positions with the half-reaction above it.

	$Cl_{2(g)} + 2\ e^- \rightleftarrows 2\ Cl^-(aq)$	+1.36
new position	$Cr_2O_7^{2-}(aq) + 14\ H^+(aq) + 6\ e^- \rightleftarrows 2\ Cr^{3+}(aq) + 7\ H_2O_{(l)}$	+1.33
	$2\ HNO_{2(aq)} + 4\ H^+(aq) + 4\ e^- \rightleftarrows N_2O_{(g)} + 3\ H_2O_{(l)}$	+1.30

Nelson CHEMISTRY

9. To become familiar with using your textbook redox table, answer textbook questions 19-26 on pages 359 and 360 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 1: Activity 3.

What I'd like to know is, where is all of this heading?
Identifying the oxidizing and reducing agents in a mixture is the first main step in predicting a redox reaction!

The redox half-reaction table is a useful tool for many electrochemical applications. In the next activity, you will use the redox table to predict more complex redox reactions and whether the reactions are spontaneous. This will eventually connect the energy aspects you studied earlier to redox reactions.

ACTIVITY 4

Predicting the Outcomes

Imagine that you are going to a dance on the weekend. You know that potentially, there will be several people to dance with, but realistically, you will have an opportunity to dance with only one partner at a time. How do you make the decision about whom to dance with? You can make only one choice. Before you decide, you must be able to see all the possibilities.

Just as there is only one partner you can choose to dance with at a time, in a mixture containing many oxidizing agents and reducing agents only two are meant for each other. They are the strongest oxidizing agent and the strongest reducing agent.

1. Why do you suppose the strongest oxidizing and reducing agents are the ones to react out of a mixture of several possible reactants?

Check your answers by turning to the Appendix, Section 1: Activity 4.

When you are asked to determine the strongest oxidizing and reducing agents, you are being asked to determine the main dance partners for a reaction. Before you can do that, you must determine all the possible oxidizing and reducing agents in the reaction mixture.

Nelson CHEMISTRY

To help you understand the process used to determine all the partners in a reaction mixture, read page 360 and the top half of page 361 in your textbook. Study the given example and make sure you note the information in Table 12.2 in the margin on page 361.

Nelson CHEMISTRY

You may find it useful to copy Table 12.2 Hints for Listing and Labelling Entities, found on page 361 of *Nelson Chemistry*, into your notebook. Add $Cr^{2+}_{(aq)}$ to your list as an ion that may act as either an oxidizing or reducing agent. Further hints for determining all the possible oxidizing and reducing agents in a reaction are given in the following chart.

Other Important Hints

- Ionic compounds in aqueous solution will be written as dissociated.

$$\text{e.g., } NaCl_{(aq)} \xrightarrow[\text{write as}]{\text{you will}} Na^+_{(aq)},\ Cl^-_{(aq)}$$

- Molecular substances in aqueous solution are not written as dissociated.

$$\text{e.g., an aqueous chlorine solution} \xrightarrow[\text{write as}]{\text{you will}} Cl_{2(aq)}$$

(**Note:** Your redox table shows chlorine as $Cl_{2(g)}$. Whether you use $Cl_{2(aq)}$ or $Cl_{2(g)}$, the half-reaction equation will be the same.)

- If you are told a strong acid is present in solution, you will write the acid in its ionized form.

$$\text{e.g., } HNO_{3(aq)} \xrightarrow[\text{write as}]{\text{you will}} H^+_{(aq)},\ NO_3^-(aq)$$

The six strong acids that fit into this category are listed on the top of the Acids and Bases table in Appendix F on page 553 of your textbook. If the acid is not strong (one of the six), you will write it as a molecular substance.

Nelson CHEMISTRY

Study the following information that gives you another example of listing the entities or species present in a potential reaction mixture. Remember, list *all* possibilities (OA/RA) for each substance present. If an ion or molecule is found on both sides of the Redox Table on page 552 of your textbook, it has the ability to act as an oxidizing agent (OA) or reducing agent (RA), or both.

Example

An acidified solution of potassium dichromate, $K_2Cr_2O_{7(aq)}$, and hydrogen peroxide liquid, $H_2O_{2(l)}$, are mixed.

Note: The boxes identify the information that is important to you.

- Acidified $\xrightarrow{\text{means}}$ $H^+_{(aq)}$ ions are present.
- $K_2Cr_2O_{7(aq)} \xrightarrow[\text{write as}]{\text{you will}}$ $K^+_{(aq)}$, $Cr_2O_7^{2-}{}_{(aq)}$.
- $H_2O_{2(l)} \xrightarrow[\text{write as}]{\text{you will}}$ $H_2O_{2(l)}$.
- Since the mixture is in aqueous conditions, $\xrightarrow[\text{include}]{\text{you}}$ $H_2O_{(l)}$.

Your list of possible participants then includes the following:

$$H^+_{(aq)},\ K^+_{(aq)},\ Cr_2O_7^{2-}{}_{(aq)},\ H_2O_{2(l)},\ H_2O_{(l)}$$

To determine whether each entity is an oxidizing agent or reducing agent, check its location on the Redox Table on page 552 of your textbook.

$H^+_{(aq)}$ → can act alone as OA

$K^+_{(aq)}$ → can act as OA

$Cr_2O_7^{2-}{}_{(aq)}$ → must be with $H^+_{(aq)}$ to act as OA

$H_2O_{2(l)}$ → can act as RA

$H_2O_{(l)}$ → can act as OA or RA (It is found on both sides of the redox table.)

Your complete identification of all entities present then looks like the following:

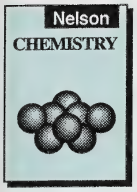
OA	OA		OA	OA
$H^+_{(aq)}$,	$Cr_2O_7^{2-}{}_{(aq)}$,	$K^+_{(aq)}$,	$H_2O_{2(l)}$,	$H_2O_{(l)}$
			RA	RA

As a practice exercise, do the following question from your textbook.

2. Answer all the parts of textbook question 29 on page 361 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 1: Activity 4.

An important ability you need to develop is predicting redox reactions in solution. This will allow you to identify the most likely reaction to occur from a mixture of several substances.



To see how this process works, read the bottom half of page 361 and all of page 362 in your textbook. Carefully study the examples given on pages 361, 362, and 364.

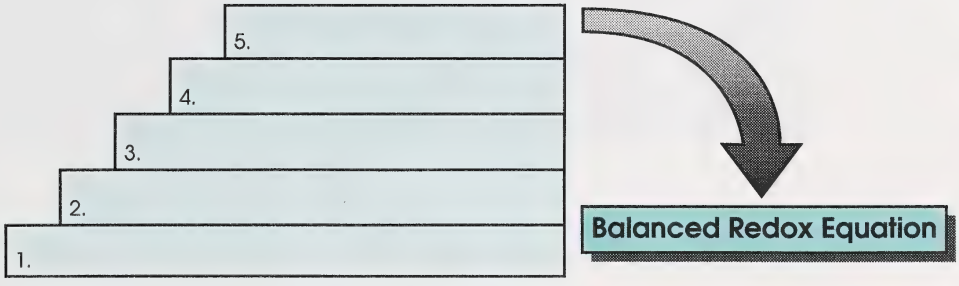
3. How do you choose the strongest oxidizing agent (SOA) and strongest reducing agent (SRA) from a list of species?

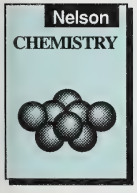
4. The half-reaction for the strongest oxidizing agent is taken directly from the redox table and is written in the same way. The strongest reducing agent half-reaction, though, is written in the reverse direction as it appears on the redox table. Why?

5. At times, why must one half-reaction or the other, or both half-reactions be multiplied by a number or factor?

6. How do you decide if the overall balanced equation is spontaneous?

7. Use the following graphic to summarize the steps involved in predicting redox equations.





8. Answer textbook question 30 (a-f) on page 364 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 1: Activity 4.

The example problem on page 364 of your textbook illustrates a practical use for the redox principles you have been studying. It combines the concept knowledge of predicting redox reactions with practical diagnostic tests, some of which you have seen before on page 537 of your textbook. Study this example again carefully before answering the next questions.

Nelson CHEMISTRY

9. Answer textbook questions 31 and 32 (a, b, c) on pages 364-365 of *Nelson Chemistry*.

10. **Lab Exercise 12D: Testing Redox Concepts**

Perform Lab Exercise 12D: Testing Redox Concepts on page 365 of your textbook. Carefully follow the given directions. Pay special attention to the applied science skills.

Science Skills

- [x] A. Initiating
- [ ] B. Collecting
- [ ] C. Organizing
- [ ] D. Analysing
- [ ] E. Synthesizing
- [ ] F. Evaluating

Check your answers by turning to the Appendix, Section 1: Activity 4.

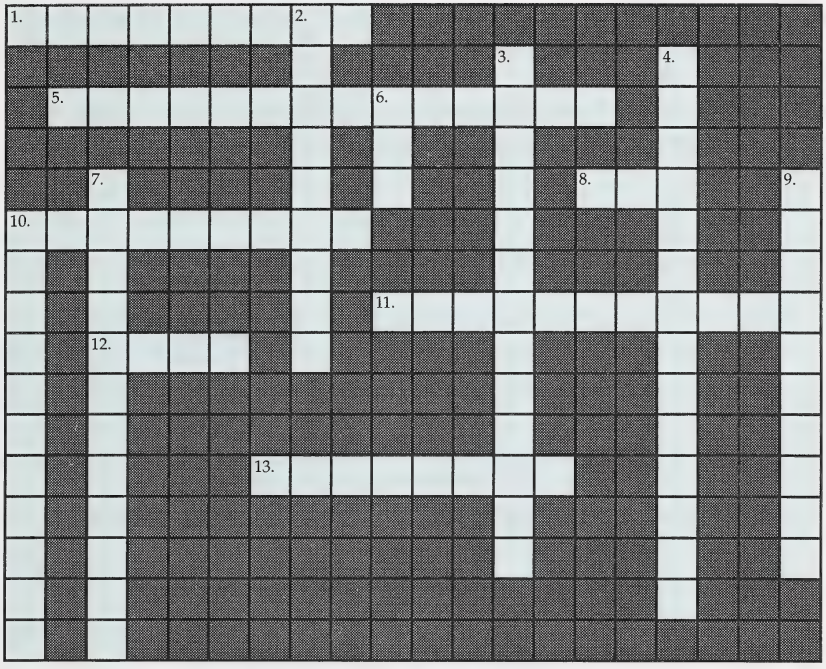
You have learned that not all redox reactions will take place automatically. To have a spontaneous redox reaction take place, the oxidizing agent must have a great attraction for the electrons and the reducing agent must have an ability to donate electrons. By making use of a redox chart of reduction half-reactions you are able to choose the dancing partners of a redox reaction and match the reacting species in such a way as to predict the spontaneity of a reaction.

Follow-up Activities

If you had difficulties understanding the concepts in the activities, it is recommended that you do the Extra Help. If you have a clear understanding of the concepts, it is recommended you do the Enrichment.

Extra Help

1. To understand the concept of electron transfer and how it relates to redox reactions, solve the following crossword puzzle.



ACROSS

1. the process by which, in a half-reaction, electrons are gained
5. chemical species which will gain electrons
8. mnemonic used to help explain a loss of electrons
10. substances being combined in a chemical reaction
11. word describing a reaction which will continue on its own
12. the product formed by the oxidation of iron
13. move from one area to another

DOWN

2. the process by which, in a half-reaction, electrons are lost
3. chemical species which will lose electrons
4. chemical reactions involving the transfer of electrons
6. mnemonic used to help explain a gain of electrons
7. a balanced equation describing one-half of an overall process
9. separate an ionic compound into ions in a solvent
10. substitution; taking the place of

2. For each half-reaction equation that follows, use the terms *reduction* or *oxidation* to denote whether the half-reaction is gaining or losing electrons.

   a. $Na_{(s)} \rightarrow Na^{+}_{(aq)} + 1\,e^{-}$

   b. $Ag^{+}_{(aq)} + 1\,e^{-} \rightarrow Ag_{(s)}$

   c. $2\,H_2O_{(l)} + 2\,e^{-} \rightarrow H_{2(g)} + 2\,OH^{-}_{(aq)}$

   d. $Ca_{(s)} \rightarrow Ca^{2+}_{(aq)} + 2\,e^{-}$

3. For each example that follows, identify the oxidizing and reducing agents, write the oxidation and reduction half-reaction equations, follow up with the net overall balanced equation, and predict the spontaneity of the reaction.

   a. A solution of $Cu(NO_3)_{2(aq)}$ is mixed with a piece of solid nickel.

   b. Acidified $KMnO_{4(aq)}$ is mixed with $CuBr_{2(aq)}$.

4. Use the results from the given chart of hypothetical elements X, Y, and Z to answer this question. (A checkmark indicates that a reaction has occurred.)

   a. List the oxidizing agents in decreasing order of strength.

   b. Write an equation for the overall net redox reaction using the strongest oxidizing agent and the strongest reducing agent.

Metals / Metal ions	$X_{(s)}$	$Y_{(s)}$	$Z_{(s)}$
$X^{2+}_{(aq)}$			✔
$Y^{2+}_{(aq)}$	✔		✔
$Z^{2+}_{(aq)}$			

Check your answers by turning to the Appendix, Section 1: Extra Help.

Nelson CHEMISTRY

If you would like further practice in working with the concepts presented in this section, answer any of questions 1-6, 10, 11-16, and 20 on pages 382-383 of your textbook. The answers to these questions can be found on pages 515-516 of *Nelson Chemistry*.

Enrichment

Do **one** of the following activities.

1. Perform the following investigation which combines some interesting redox reactions with some built-in diagnostic tests. Pay special attention to the required components, safety aspects, and applied science skills.

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☑ E. Synthesizing
- ☐ F. Evaluating

Investigation: Colour Changes in Redox Reactions

Problem

What redox reactions will occur when galvanized nails are mixed with solutions of household bleach and iodine?

Materials

- galvanized tacks or small nails (available from a hardware store)
- household bleach
- tincture of iodine (available from a drugstore or pharmacy)
- white vinegar
- transparent glass or cup

Procedure

Caution

- Be sure to use care and caution when handling the chemicals involved.
- Take a glass cup and place enough galvanized nails into it to just cover the bottom of the glass.
- Pour enough iodine into the glass so that it just barely covers the nails.
- Wait until the colour of the iodine solution changes; then slowly pour enough bleach into the solution to cause a second colour change.
- To clear any precipitate that may form, pour in some vinegar.

Analysis

Assuming the galvanized nails have a coating of solid zinc, the iodine tincture is elemental iodine, and the bleach contains elemental chlorine, propose half-reaction equations that would explain what happened with each colour change.

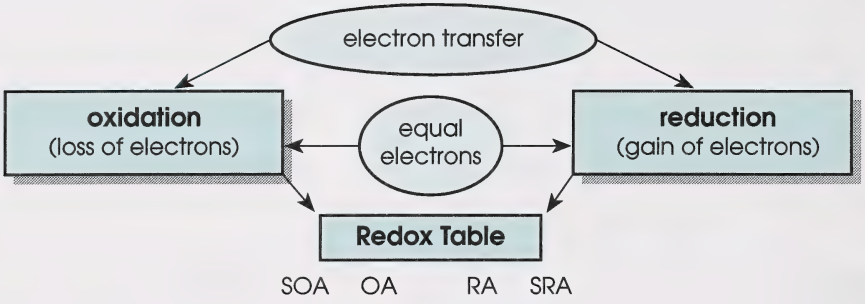
2. It is known that the aging process causes a deterioration both physically and mentally. Your body is composed of millions of cells, each group of which is responsible for specific functions. It is also known that oxygen is an oxidizing agent. Briefly explain if it would be conceivable that the air you breathe could be in part responsible for aging.

3. a. Corrosion is the erosion and disintegration of a material resulting from a chemical reaction. Rusting is the corroding of iron to form reddish brown oxides of iron. Propose two half-reactions from a redox chart, and the overall reaction, which would show what happens when an iron bar is exposed to the air.

   b. Visit an autobody shop or talk to a chemical engineer to discover what steps can be and are taken to prevent or slow down the corrosion of iron in today's vehicles.

Check your answers by turning to the Appendix, Section 1: Enrichment.

Conclusion



You have discovered that all electrochemical reactions involve a loss and gain of electrons, and therefore are redox reactions. There must be both a gain and a loss of electrons for a redox reaction to take place. Evidence from many different types of redox reactions has allowed for the formation of a redox chart, which enables you to predict the most likely outcome of reacting species. The next section will give you the opportunity to apply these redox principles in practical situations.

40

Section 1 Assignment: Electron Transfer and Reactions

Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

Chemistry 30 – Module 3 Section 1 Assignment Page # Name and ID#

Be sure to write legibly. Leave a wide left margin and number all of your pages.

1. Using your periodic table of the elements, indicate whether each of the following substances could act as an oxidizing agent (OA), a reducing agent (RA), or possibly both (OA/RA). **(4 marks)**

   a. $Ba^{2+}_{(aq)}$ b. $Co^{2+}_{(aq)}$ c. $Se_{(s)}$ d. $Cs_{(s)}$

Science Skills
- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☑ E. Synthesizing
- ☐ F. Evaluating

2. A student completed an investigation to identify some unknown metals, which were labelled D, Z, A, and X. The choices for the metals were Zn, Pb, Fe, and Ag. The following chart gives results of reactions that occurred between the metal ions and metals that were placed in the solutions. A positive result for a reaction is indicated by the word *yes*.

Metal ions / Metals	$D^{2+}_{(aq)}$	$Z^{2+}_{(aq)}$	$A^{2+}_{(aq)}$	$X^{+}_{(aq)}$
$D_{(s)}$				yes
$Z_{(s)}$	yes			yes
$A_{(s)}$	yes	yes		yes
$X_{(s)}$				

   a. From the results in the table, create a redox table of reduction half-reactions. List the oxidizing agents in decreasing order of strength. **(4 marks)**

   b. Match each letter of the unknown metals with its correct metal symbol. **(2 marks)**

Response Page Setup

Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

3. For each of the following, indicate whether the equation represents a reduction half-reaction or an oxidation half-reduction. **(5 marks)**

   a. $2\ H_2O_{(l)} + 2\ e^- \rightarrow H_{2(g)} + 2\ OH^-_{(aq)}$

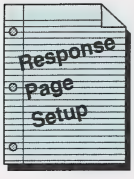
   b. $S_{(s)} + 3\ H_2O_{(l)} \rightarrow H_2SO_{3(aq)} + 4\ H^+_{(aq)} + 4\ e^-$

   c. $As_2O_{3(s)} + 6\ H^+_{(aq)} + 6\ e^- \rightarrow 2\ As_{(s)} + 3\ H_2O_{(l)}$

   d. $Fe^{3+} + 3\ e^- \rightarrow Fe_{(s)}$

   e. $Pb_{(s)} + SO_4{}^{2-}{}_{(aq)} \rightarrow PbSO_{4(s)} + 2\ e^-$

4. Explain historically, in one-quarter page or less, what the terms *oxidation* and *reduction* originally meant (give the empirical definitions). **(2 marks)**



5. Could you successfully store a solution of acidified potassium permanganate in a tin can? Explain. Support your answer to this question with chemical equations. **(4 marks)**

6. For each of the following situations, write out the most likely redox equation. Follow the five-step procedure for predicting redox reactions. Show all your steps.

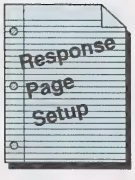
   a. Zinc metal is placed into a solution of $AuCl_{3(aq)}$. **(3 marks)**

   b. Chlorine gas and hydrogen gas are simultaneously bubbled into a sodium hydroxide solution. **(3 marks)**

   c. A perchloric acid solution is slowly added to hydrogen peroxide, $H_2O_{2(l)}$. **(3 marks)**

   d. A fountain pen dipped into hydrochloric acid is used to etch a student's name onto a piece of copper foil. **(3 marks)**

7. Did the pen in question 6. d. successfully etch the copper foil? Explain. **(2 marks)**



8. Predict the products of a reaction between a small piece of sodium metal and water. Give three diagnostic tests that could be used to identify the products that were formed. Support your answer for this question with chemical equations. **(5 marks)**

SECTION 2

Redox Applications

You pick up a newspaper and read that a provincial government agency has determined that a river water supply is above all the standards set for safe use. You later hear a television news report of protesters quoted as saying that they have data which proves a pulp and paper mill is polluting this river. How do the provincial agency or the protest group know whether or not the water contains pollutants? Can they actually see particles floating around in it, or perhaps the water gives off a bad odour? Each group has read laboratory reports. When chemists analyse a solution, such as a water sample, they may use redox principles and a technique called titration. When they take the data from the titration and use mathematics to determine the actual amounts of what is present in the water they are using stoichiometric methods to quantify their work.

In this section you will apply the redox concepts worked with in Section 1 along with the methods of stoichiometry and titration analysis to analyse various solutions. You will discover how to identify a redox reaction based on the assignment of oxidation numbers to the reactants and products. Finally, you will demonstrate the use of oxidation numbers and half-reactions in the balancing of redox equations.

ACTIVITY 1

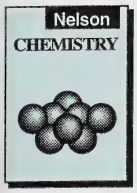
Redox Stoichiometry

If you were given the task of analysing some river water, would you have to survey the entire river? Or, if you wanted to discover the contents of cement, would you need a whole truckload?

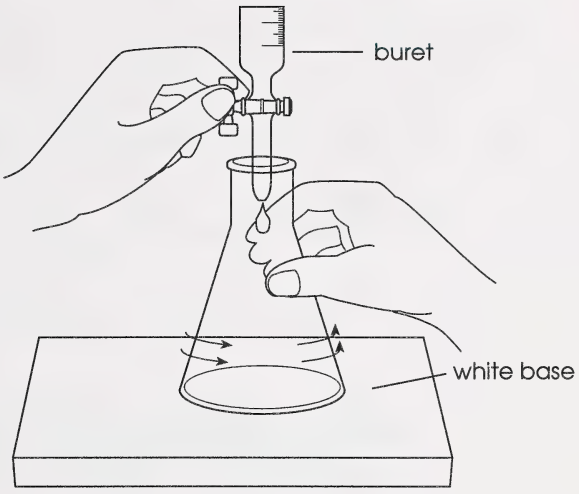
Often analyses such as these are done on a small sample of material, but precise results are still obtained. Lab technicians who analyse, for example, samples of water sent to them, play a Sherlock Holmes type of game to determine the amounts of certain substances present in the water. The identification of chemical species can be done by several different techniques. The choice of technique can often depend on what is thought to be in the water. One analysis technique often used is called titration analysis.

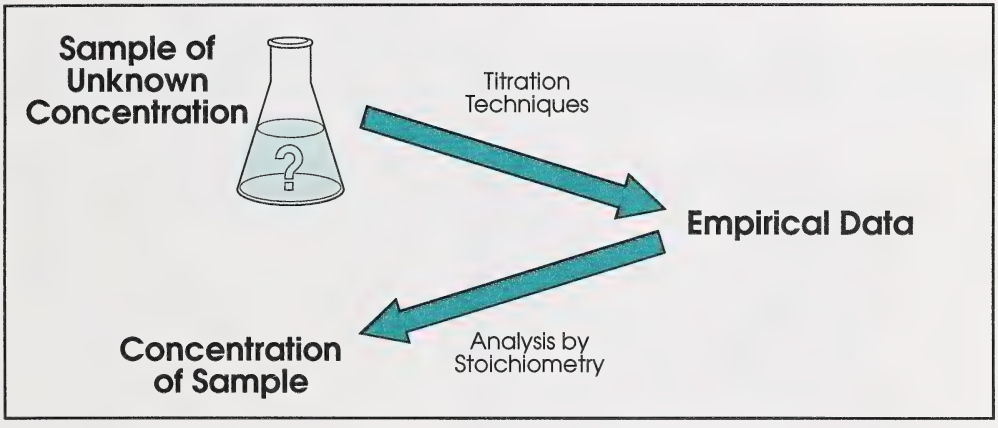
1. You were introduced to titration analysis in Chemistry 20. What is involved in titration analysis?

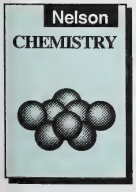
Check your answers by turning to the Appendix, Section 2: Activity 1.



In Chemistry 20 you discovered how titration techniques can be used when dealing with acid-base reactions. The same techniques can be used when working with redox reactions. If you have trouble remembering what was involved in the titration techniques you learned in Chemistry 20, you can review pages 190-191 and 536-537 in your textbook to refresh your memory.

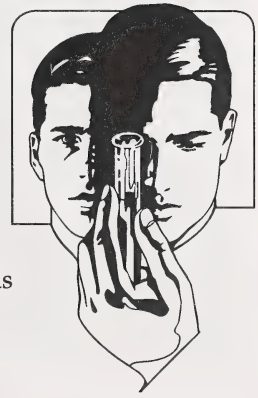






As you also discovered in Chemistry 20, stoichiometry can be used in conjunction with titration techniques to analyse an unknown substance. The stoichiometric problem-solving techniques used in Chemistry 20 will also be used in this course. You may wish to review these techniques from your Chemistry 20 materials.

To see how stoichiometry and titration principles can apply to redox reactions, read the bottom of page 365 and the top two-thirds of page 366 in *Nelson Chemistry*.



2. List three industrial or laboratory applications of redox stoichiometry.

3. List the missing terms in the following passage.

   In a titration one chemical substance called the ____(a.)____, is slowly added to the ____(b.)____ being tested. The process is continued until an abrupt change in the solution property, the ____(c.)____, is reached. In redox reactions this is usually a ____(d.)____. This abrupt change usually signals the point at which chemically equivalent amounts of reactants are present. This point is called the ____(e.)____ point.

Check your answers by turning to the Appendix, Section 2: Activity 1.

It is important that you do not mix up the terms *equal* and *equivalence point*. The amounts of each reacting species at an equivalence point may or may not be equal. Only the balanced equation for the reaction can tell you what amounts of each species must be present at the equivalence point. To get a better understanding of this concept do the following question.

4. For each of the following balanced redox reactions determine the amount of oxidizing agent, by mole, that would be necessary to react with the given amount of reducing agent, to reach an equivalence point.

   a. $2\ Sr_{(s)} + O_{2(g)} \rightarrow 2\ SrO_{(s)}$

RA	OA
4.0 mol	? mol

   b. $3\ Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$

RA	OA
6.0 mol	? mol

   c. $Zn_{(s)} + I_{2(s)} \rightarrow ZnI_{2(s)}$

RA	OA
2.5 mol	? mol

   d. $4\ Fe_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Fe_2O_{3(s)}$

RA	OA
12.0 mol	? mol

   e. $Si_{(s)} + 2\ F_{2(g)} \rightarrow SiF_{4(s)}$

RA	OA
3.5 mol	? mol

Check your answers by turning to the Appendix, Section 2: Activity 1.

You learned in Chemistry 20 that observable evidence for the endpoints of acid-base titrations were colour changes of indicators added to the solution. For redox titrations, you will also use a colour change as the evidence for an endpoint. However, the colour change will happen to one of the reactants itself, not an added indicator. Many substances can be used as redox reaction titrants because of their distinct colour changes.

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5. For each of the following three reduction half-reactions, identify the species which have a distinct colour. Indicate their appropriate colours by using the Ion Colours chart on the inside back cover of your textbook.

   a. $MnO_4^-{}_{(aq)} + 8\,H^+{}_{(aq)} + 5\,e^- \rightarrow Mn^{2+}{}_{(aq)} + 4\,H_2O_{(l)}$

   b. $Cr_2O_7^{2-}{}_{(aq)} + 14\,H^+{}_{(aq)} + 6\,e^- \rightarrow 2\,Cr^{3+}{}_{(aq)} + 7\,H_2O_{(l)}$

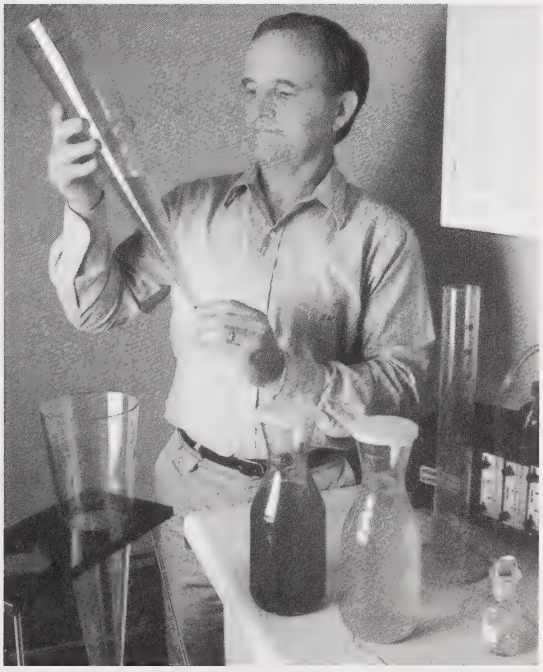
   c. $Fe^{3+}{}_{(aq)} + 1\,e^- \rightarrow Fe^{2+}{}_{(aq)}$

6. Explain the difference between an endpoint and an equivalence point.

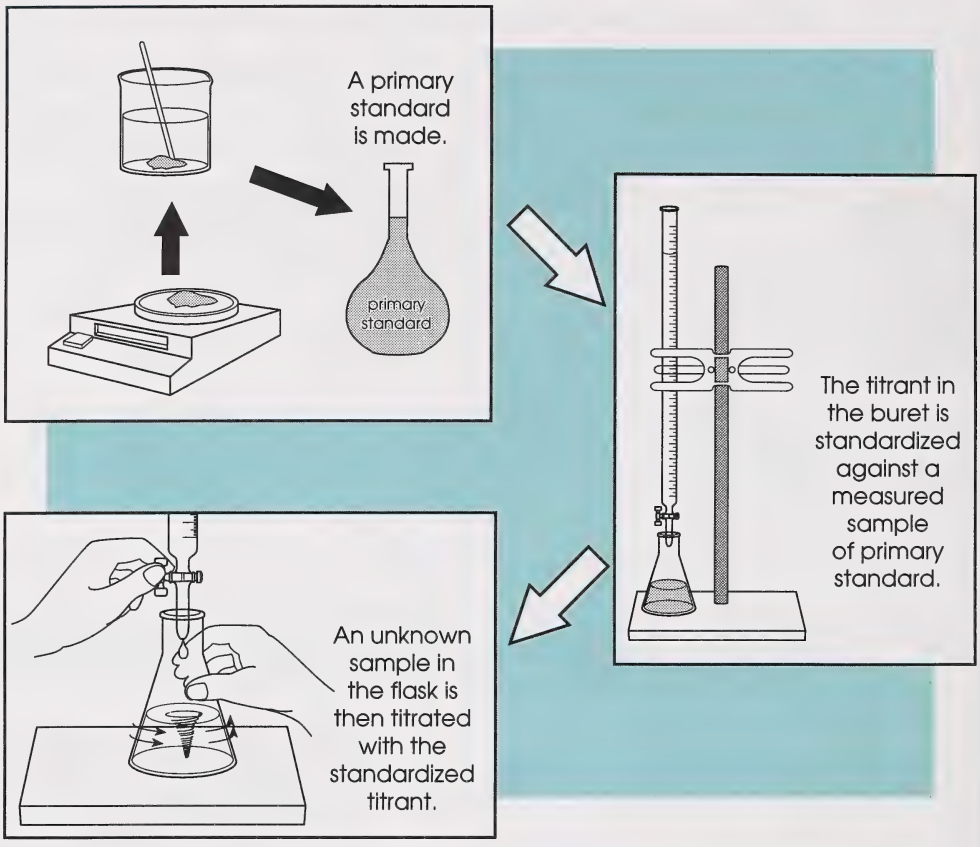
Check your answers by turning to the Appendix, Section 2: Activity 1.

Put yourself in the place of an industrial chemist. Can you imagine giving someone the results of a water sample test based on data of a titrant that was sloppily made, or one that you weren't sure of in terms of its accurate concentration? It would throw off the results you would have obtained for your sample. An entire industrial complex could be shut down based on your incorrect results.

When a solution used as a titrant has been accurately made and its concentration is precisely known, it is said to be standardized. The following flow chart shows one example of the process that is used in preparing these standardized solutions.



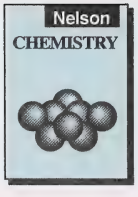
NASA


A primary standard is made.
primary standard
The titrant in the buret is standardized against a measured sample of primary standard.
An unknown sample in the flask is then titrated with the standardized titrant.

primary standard – a stable chemical that can be used in a standard solution

A **primary standard** is a chemical that could be used as the titrant or could be used to find the precise concentration of a titrant. It is stable enough to not undergo any reaction with anything, including air, before you want it to.

Two poor substances to use as primary standards would be potassium permanganate and sodium hydroxide. Potassium permanganate, although great to use because of its colour change, is such a strong oxidizing agent that it oxidizes almost anything it comes in contact with. It could actually react to some degree with the water and air before you let it come in contact with the desired reactant. Sodium hydroxide (in its solid state), on the other hand, absorbs water from the air so easily that you can never get an accurate mass reading for it. Using a primary standard to standardize either of these in solution form before using them as a titrant becomes very important to chemists.



To better understand this idea, review page 366 of your textbook. Then read through Lab Exercise 12E: Standardizing Potassium Permanganate (pages 366-367) and answer the following questions, based on the lab exercise.

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Lab Exercise 12E: Standardizing Potassium Permanganate

7. Why is potassium permanganate not used as a primary standard?

8. What chemical substance is going to be used as the primary standard?

9. How can you confirm that the balanced equation for the redox reaction involved in the standardization of potassium permanganate is spontaneous?

10. Four trials were carried out, but only three trial results were used in calculating the concentration of the permanganate ion solution. Explain why.

11. What colour change indicated the endpoint of the reaction?

12. The calculation of the $MnO_4^-{}_{(aq)}$ concentration in your textbook is written in step form. Rewrite the calculation in one-line form and show how the units cancel (as you did in Chemistry 20).

13. a. Considering the predicted concentration of the potassium permanganate solution, and the standardized concentration, what percent error (percent difference) exists between the two results of the concentration for the permanganate ion?

    b. Would this mean that permanganate ion solutions should be standardized? Explain.

Check your answers by turning to the Appendix, Section 2: Activity 1.

Here is an important point to note:

You will be doing many different stoichiometry problems as you work through the Chemistry 30 program. Whether the problems deal with aspects of energy, redox, or acids and bases, it is important to remember that all stoichiometry problems follow the same general pattern. The difference is that you may have to write a redox reaction in one case or an acid-base reaction in another; the calculation steps are the same.

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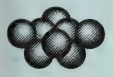


There are many redox applications you can choose for studying the usefulness of redox stoichiometry and titration procedures. The next three lab exercises from your textbook explore a variety of these applications. Read pages 368-369; then answer the following questions.

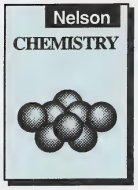
Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

14. a. Do the analysis of the report for Lab Exercise 12F: Analysing for Tin on page 368 of your textbook.

    b. Complete the analysis for Lab Exercise 12G: Analysis of Chromium in Steel on page 368 of your textbook.

Check your answers by turning to the Appendix, Section 2: Activity 1.



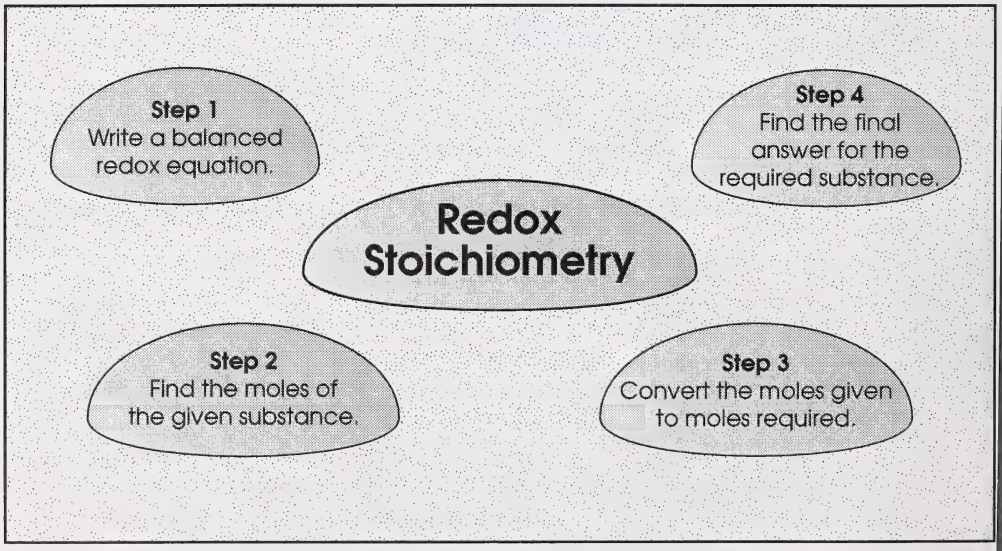
In each of the preceding lab exercises you generated a redox reaction and then performed a stoichiometric calculation based on that reaction and the evidence provided. The next lab exercise is similar, but, in addition, you are asked to do an evaluation of the process involved.

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☑ E. Synthesizing
- ☑ F. Evaluating

15. Complete the analysis for Lab Exercise 12H: Analysing for Iron on page 369 of your textbook. Then consider the prediction and a percent error calculation to complete the evaluation.

Check your answers by turning to the Appendix, Section 2: Activity 1.


Step 1
Write a balanced redox equation.
Step 4
Find the final answer for the required substance.
Redox Stoichiometry
Step 2
Find the moles of the given substance.
Step 3
Convert the moles given to moles required.

You have been introduced to the concept of standardizing a solution for titration purposes. In the following investigation you will standardize a solution and then use that solution to titrate a hydrogen peroxide solution.

Investigation 12.4: Analysis of a Hydrogen Peroxide Solution

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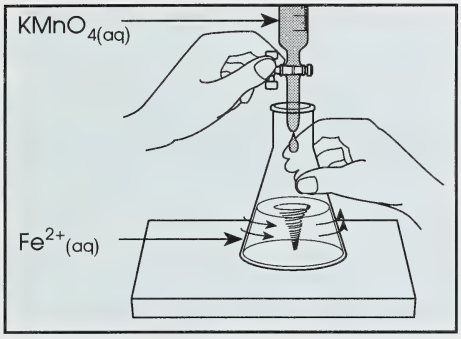
Carefully follow the directions given for this investigation on pages 371-372 of your textbook. Pay special attention to the required components, safety aspects, and applied science skills.

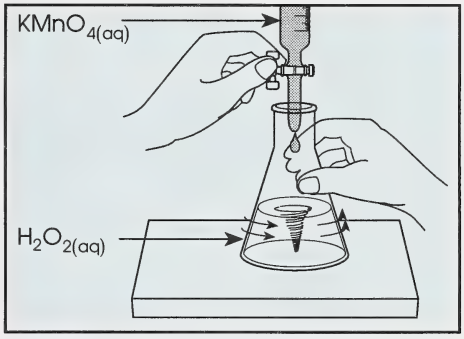
Guidelines

- The procedure to this lab will be done in two stages: standardization and titration.

Stage 1: Standardization



Stage 2: Titration



- The $KMnO_{4(aq)}$ is standardized by using $FeSO_4 \bullet (NH_4)_2SO_4 \bullet 6H_2O_{(aq)}$ which is prepared as a primary standard. (Only the $Fe^{2+}{}_{(aq)}$ from this compound is important to you. It is the ion that participates in the reaction.)

- The prepared $FeSO_4 \bullet (NH_4)_2SO_4 \bullet 6H_2O_{(aq)}$ must be acidified with $H_2SO_{4(aq)}$.

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- If you have forgotten how to use a pipet, review the instructions given on pages 532 and 533 of your textbook.

- You may also wish to review the procedures for preparing standard solutions and titrating on pages 535 and 536 of your textbook.

16. Make a prediction as to the percent concentration of hydrogen peroxide in the consumer product you will be testing.

17. Perform Step 1 of the procedure (page 371) which asks you to do a pre-lab calculation.

Check your answers by turning to the Appendix, Section 2: Activity 1.

PATHWAYS

If you have access to supervised laboratory facilities, do Part A.
If you do not have access to laboratory facilities, do Part B.

Part A

Science Skills

- ☑ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

18. Perform Steps 2-5 of the procedure (textbook page 372). Record your evidence in a table similar to the one on the bottom of page 370 of your textbook.

19. Perform Steps 6-9 of the procedure (textbook page 372). Record your evidence in a table similar to the one used for the standardization in the preceding question.

Check your answers by turning to the Appendix, Section 2: Activity 1.

As your analysis, answer the questions after Part B.

End of Part A

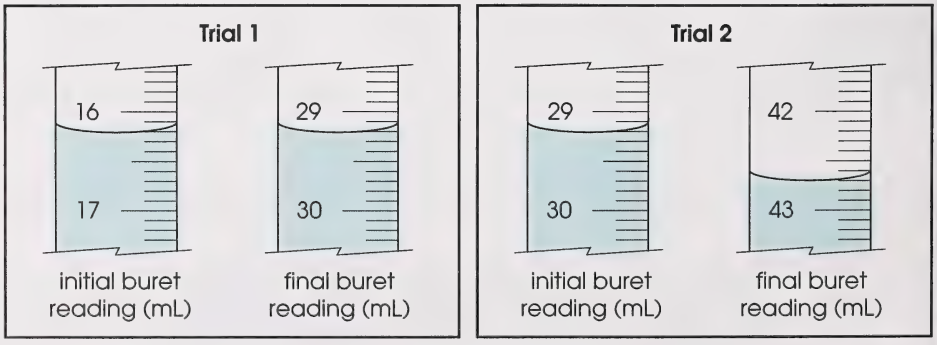
Part B

Science Skills

- ☑ A. Initiating
- ☐ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

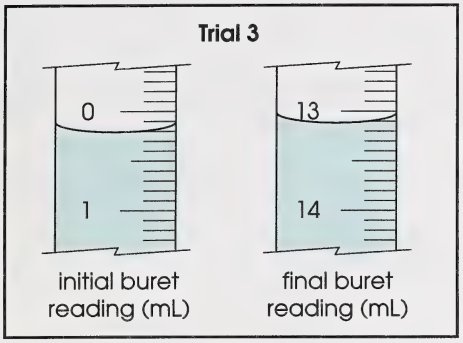
The nature of the chemicals involved prevents you from actually performing this investigation. You will obtain evidence by completing the following two questions.

20. Read Steps 2-5 of the procedure (on page 372 of your textbook). The steps describe the standardization of the $KMnO_{4(aq)}$ solution. Obtain the evidence for these steps by reading the buret volumes (in millilitres) for each trial as given in the following diagrams. Record your evidence in a table similar to the one shown on page 370 of your textbook.



Endpoint colour – pink

Endpoint colour – red



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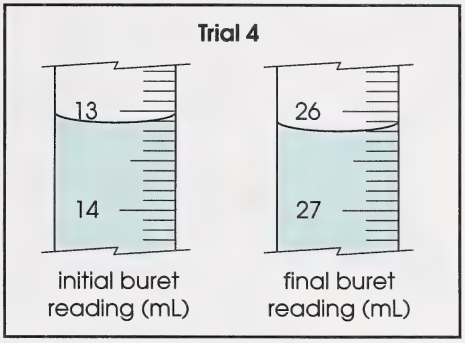
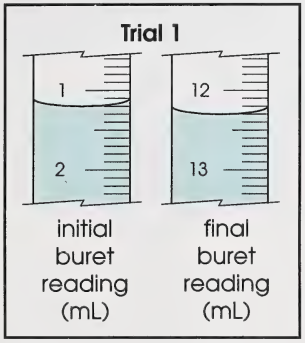


Endpoint colour – pink

21. Read Steps 6-9 of the procedure (textbook page 372). These steps describe the titration of the $H_2O_{2(aq)}$ solution. Obtain the evidence for these steps by reading the buret volumes (mL) for each trial as given in the following diagrams. Record your evidence in a table similar to the one used in the preceding question.



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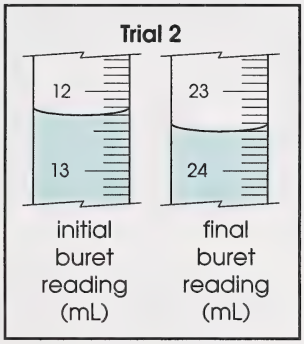
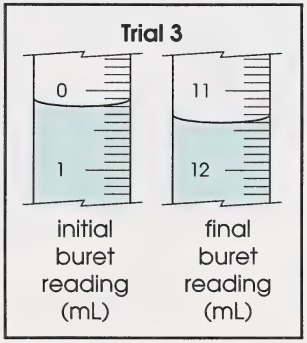


Endpoint colour – pink



Endpoint colour – pink

Check your answers by turning to the Appendix, Section 2: Activity 1.

Do the following questions to complete your analysis and evaluation.

End of Part B

Analysis

22. Why would the solution that the $FeSO_4 \bullet (NH_4)_2SO_4 \bullet 6H_2O_{(s)}$ was dissolved in need to contain $H_2SO_{4(aq)}$?

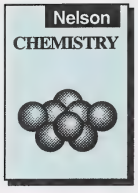
23. What was the calculated average of the three net volume readings for Stage 1 – Standardization?

24. Calculate the concentration of the $KMnO_{4(aq)}$ from Stage 1.

25. What was the calculated average of the three net volume readings for Stage 2 – Titration?

26. Calculate the concentration of the $H_2O_{2(aq)}$.

27. According to manufacturer's labels, the $H_2O_{2(aq)}$ bought in a store is 3% by concentration. Your sample of $H_2O_{2(aq)}$ was diluted fortyfold, meaning that forty times your calculated concentration will give you the original concentration of the $H_2O_{2(aq)}$. Calculate the original molar concentration of your sample of $H_2O_{2(aq)}$.



28. a. Construct a graph on graph paper of the percent concentration and molar concentration of $H_2O_{2(aq)}$ using the data given in Table 12.3 on page 371 of your textbook. Use the percent concentration as your manipulated (independent) variable.

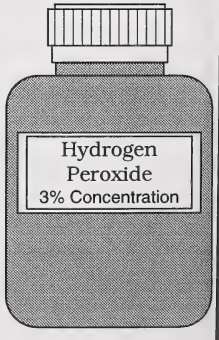
    b. Using the calculated molar concentration of $H_2O_{2(aq)}$ from the preceding question, interpolate the percent concentration of the $H_2O_{2(aq)}$ from your graph.

29. Evaluate your investigation by completing the following:

    a. Calculate a percent difference (percent error) for the determined concentration of $H_2O_{2(aq)}$.

    b. Verify the prediction.

    c. Verify the titration design.

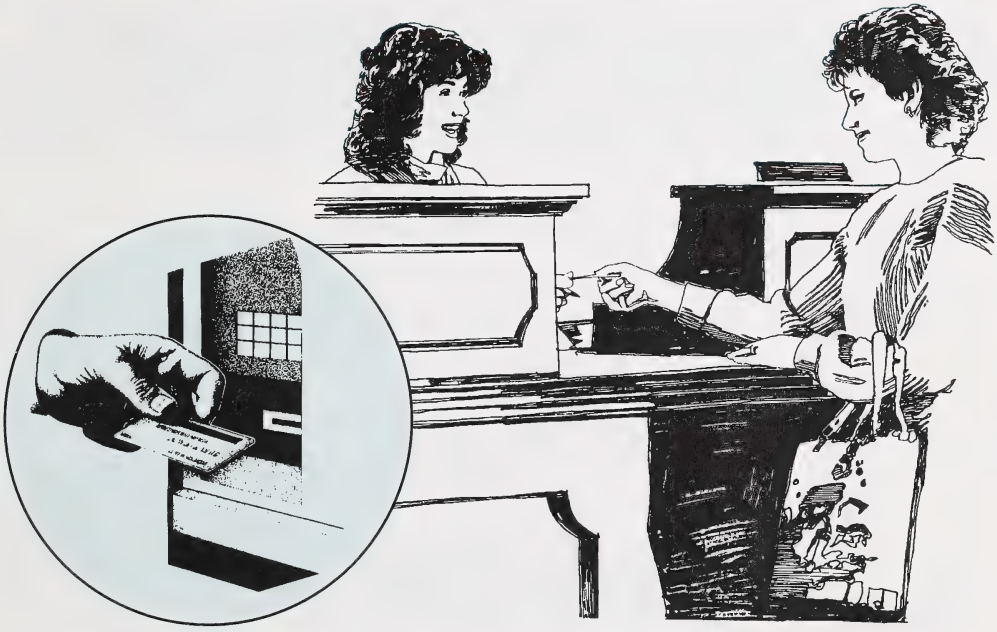


Check your answers by turning to the Appendix, Section 2: Activity 1.

As you have seen, the techniques used to study redox principles are similar to those needed to study many other aspects of chemistry.

ACTIVITY

2 Oxidation States and Numbers



What is the first thing that is assigned to you when you open a bank account? What about when you register as a student of a new school or apply for a bank card? The answer is a number.

1. Why do you think institutions often acknowledge you as a number?

oxidation number – a positive or negative number representing the oxidation state of an atom

Check your answers by turning to the Appendix, Section 2: Activity 2.

Chemists also use numbers as a way of keeping track of how chemical species can change in a reaction by losing or gaining electrons. Just as you may have a bank account number or a social insurance number, a chemical species would have an **oxidation number**.

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To learn more about oxidation numbers and their importance to redox concepts, read page 372 and the top of page 373 in your textbook.

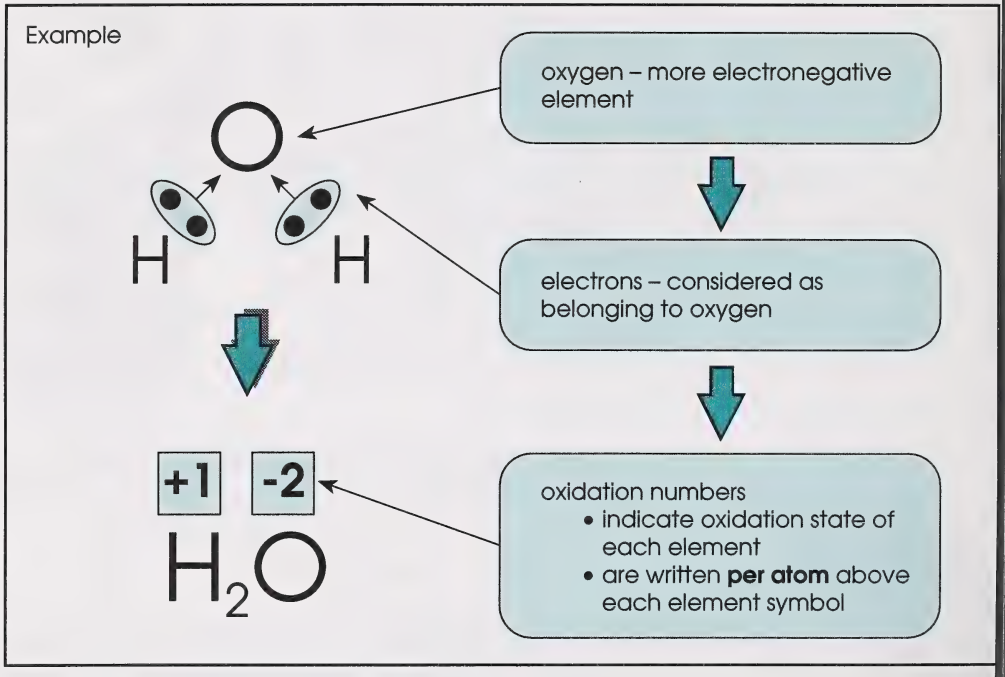
2. Why did chemists develop a method of electron bookkeeping?

3. a. What does it mean when you refer to the oxidation state of an atom?

   b. Is the oxidation state of an atom a real positive or negative charge?

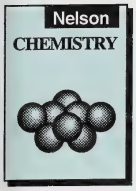
4. How are oxidation numbers distinguished from actual ion charges?

Check your answers by turning to the Appendix, Section 2: Activity 2.



Why would you want to assign oxidation numbers to elements? What purpose do they serve? Oxidation numbers enable you to perform three important tasks. They enable you to do the following:

1. Identify whether or not a reaction is a redox reaction.
2. Identify the oxidizing and reducing agents in a redox reaction.
3. Balance both simple and complicated redox reactions.



The rules and guidelines for assigning oxidation numbers are outlined on pages 373 and 374 of your textbook. Study the information given in Table 12.4, the margin information, and the examples very carefully.

5. Reproduce the following table. Fill in the missing information.

ASSIGNING COMMON OXIDATION NUMBERS

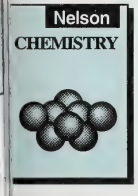
Atom or Ion	Oxidation Number	Examples
hydrogen in compounds		
oxygen in compounds		
		Mg is 0; N in N_2 is 0
	same size as charge on the ion	

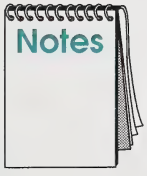
6. Describe the exceptional cases when assigning oxidation numbers for oxygen and hydrogen.

7. For a substance such as NH_3, describe how you assign the oxidation number for *N*.

8. For an ion such as $S_2O_3^{2-}$, describe how you assign the oxidation number for *S*.

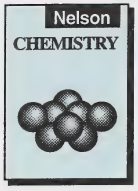
Check your answers by turning to the Appendix, Section 2: Activity 2.





You would probably find it very helpful to copy the information from the Summary: Determining Oxidation Numbers from pages 373 and 374 in your textbook onto your own paper. Along with the summary, the following two guidelines will help you further.

- The oxidation number of all Group 1 (IA) elements is +1, which is the same number as their charge.
- The oxidation number of all Group 2 (IIA) elements is +2, which is the same number as their charge.



9. Complete textbook questions 39-41 on page 374 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 2: Activity 2.

How would you assign the oxidation number for *Mn* in $MnSO_4$? It's a little more difficult, isn't it? The problem with this example, $MnSO_4$, is that two elements, Mn and S, do not have common oxidation numbers when in compounds. If you were given the hint, *think of $MnSO_4$ as an ionic compound*, would that help?

Since $MnSO_4$ is ionic, think of the two ions that compose it, Mn^{2+} and SO_4^{2-}. If you consider each ion **individually**, it makes the job of assigning oxidation numbers much easier.

Mn^{2+} → Mn^{2+} is a monatomic ion; therefore its oxidation number has the same number as its ion charge, +2.

↓

$\boxed{+2}$

Mn^{2+}

SO_4^{2-} → Now sulphur's oxidation number can be determined. You know oxygen's oxidation number is –2 and the charge on the polyatomic ion is $2-$.

↓

$1(x) + 4(-2) = -2$

$x = -2 + 8$

$x = +6$

$\boxed{+6}\ \boxed{-2}$

$S\ O_4^{2-}$

Putting all the information together, you get $\overset{\boxed{+2}\ \boxed{+6}\ \boxed{-2}}{Mn\ S\ O_4}$.

10. Determine the oxidation numbers of each element in the following substances.

a. $FeSO_3$

b. $AgNO_3$

c. $Pb_3(PO_4)_2$

d. $Al(H_2PO_4)_3$

e. $K_2Cr_2O_7$

Check your answers by turning to the Appendix, Section 2: Activity 2.

Now that you have discovered how to assign oxidation numbers you will see how oxidation numbers apply to redox reactions. The main idea here is that when a chemical substance changes its oxidation number in a reaction it will be losing or gaining electrons.

Example

- Reduction half-reaction equation (with oxidation numbers) for silver ions

$$\overset{+1}{Ag^{+}}_{(aq)} + 1e^{-} \rightarrow \overset{0}{Ag}_{(s)}$$

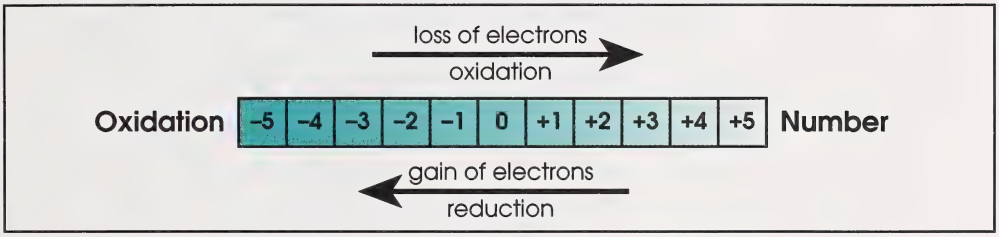
Note: The oxidation number of silver changes from +1 to 0, a **gain** of one electron.

- Oxidation half-reaction equation (with oxidation numbers) for zinc.

$$\overset{0}{Zn}_{(s)} \rightarrow \overset{+2}{Zn^{2+}}_{(aq)} + 2e^{-}$$

Note: The oxidation number of zinc changes from 0 to +2, a **loss** of two electrons.

The examples illustrate the point that, since electrons are *negatively* charged, if a substance **loses** electrons, it will show an **increase** in oxidation number. Conversely, if a substance **gains** electrons, it will show a **decrease** in oxidation number.



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For more information on the connection between oxidation numbers and redox reactions, read pages 374 and 375 in your textbook. Study the examples and the margin information carefully.

11. Answer textbook question 43 on page 375 of *Nelson Chemistry*. In addition to the given directions, connect the changing elements with arrows and indicate the process involved, as shown in the examples on page 375.

Check your answers by turning to the Appendix, Section 2: Activity 2.

From your studies in Science 10 and Chemistry 20, you should be familiar with several types of reactions. However, until this point, you may not have realized that many of those reactions were redox reactions. A wide variety of redox reactions occur in both living and nonliving systems.

Consider the following example of the photosynthesis reaction, a redox reaction that occurs in plants.

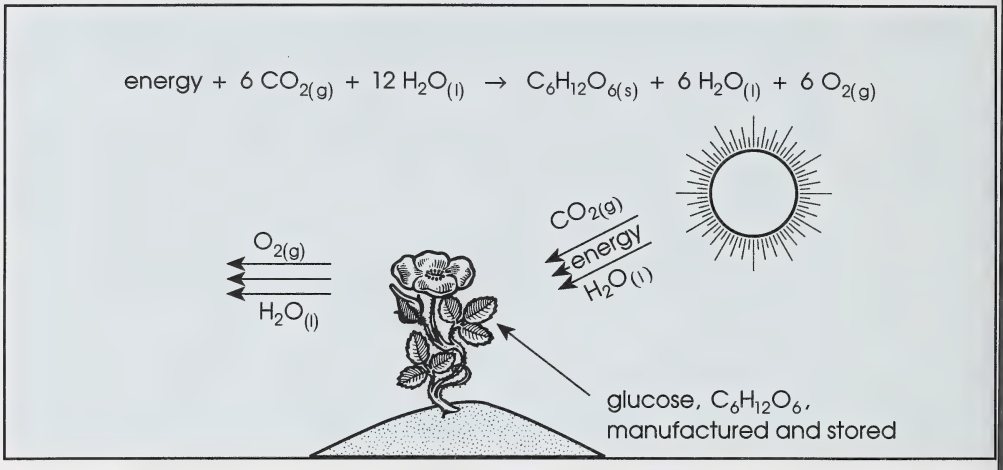


FIGURE 3.2: Photosynthesis

12. Use Figure 3.2 to answer this question.

    a. Justify that the net photosynthesis reaction is a redox reaction by assigning oxidation numbers to all species, labelling the oxidizing and reducing agents, and labelling the reduction and oxidation changes with arrows.

    b. If you were to write a reduction half-reaction equation for the photosynthesis reaction, what reactant(s) and product(s) would be in the half-reaction equation from the overall equation?

    c. What reactant(s) and product(s) would be in the oxidation half-reaction equation from the overall equation?

    d. Tapping your knowledge from Modules 1 and 2, which examine energy, calculate the enthalpy change (ΔH°_r) of the photosynthesis reaction. Where does the energy come from to carry out this reaction?

Check your answers by turning to the Appendix, Section 2: Activity 2.

One redox reaction important to many people is the combustion of natural gas (methane, CH_4) in their homes. If your house is heated with natural gas, you rely heavily on this redox reaction in cold winter months.

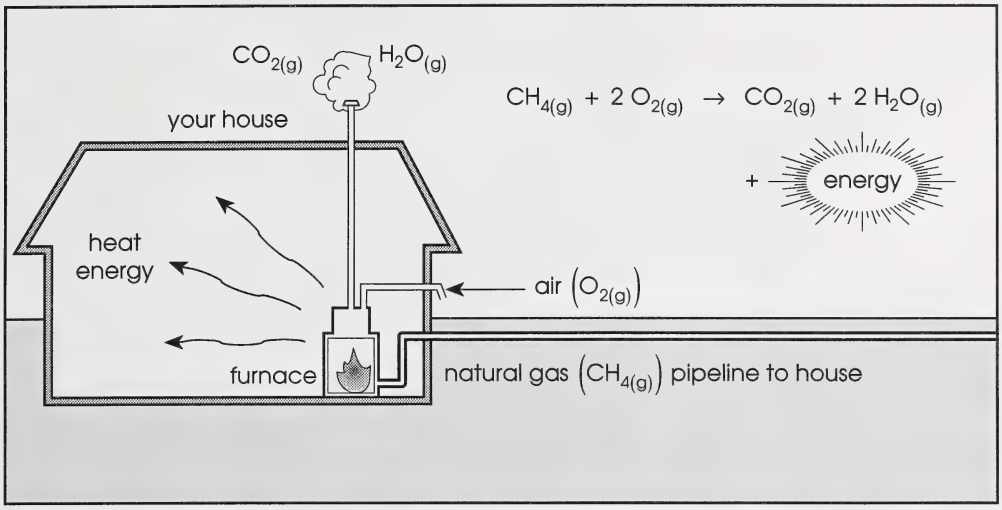


FIGURE 3.3: Methane Combustion

13. Use Figure 3.3 to answer this question.

    a. Justify that the methane combustion is a redox reaction by assigning oxidation numbers to all species, labelling the oxidizing and reducing agents, and labelling the reduction and oxidation changes with arrows.

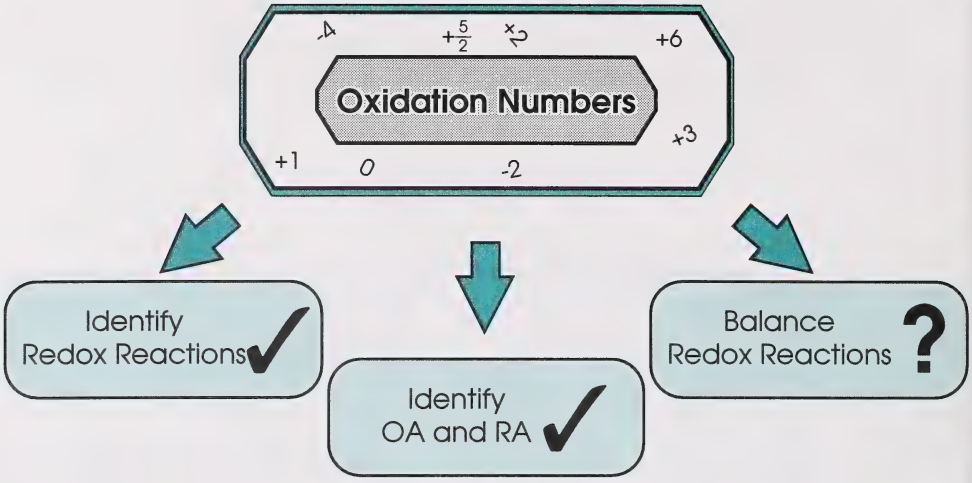
    b. If you were to write a reduction half-reaction equation for the combustion of methane, what reactant(s) and product(s) would be in the half-reaction equation from the overall equation?

    c. What reactant(s) and product(s) from the overall equation would be in the oxidation half-reaction equation?

    d. Calculate the enthalpy change for the methane combustion reaction.

14. Compare the living system photosynthesis reaction with the nonliving system methane combustion reaction. List two ways they are similar. List two ways they are different. What is the connection between the two reactions?

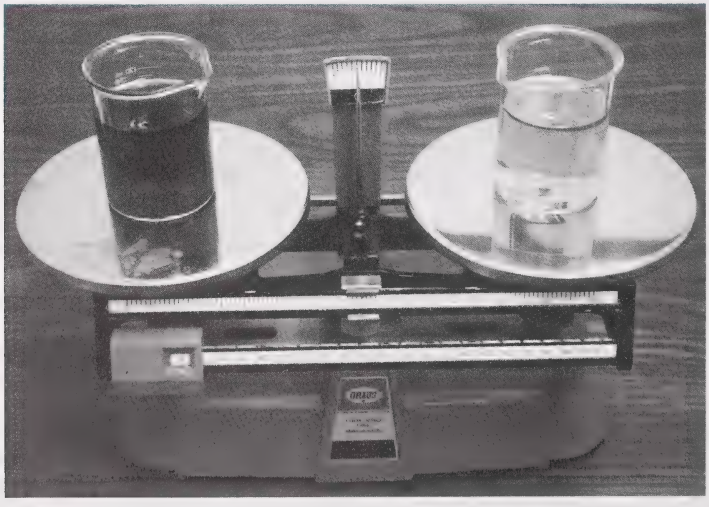
Check your answers by turning to the Appendix, Section 2: Activity 2.



Just as numbers are important in keeping track of your bank account, you have seen how oxidation numbers help to keep track of the electrons in a redox reaction. You will see an extension of the role of oxidation numbers in the next activity as you use them to balance redox equations.

ACTIVITY 3

Redox Equations in Balance



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You have been diligently placing your earnings from a part-time job into a bank account. At various times you withdraw a little money for entertainment or clothing expenses. You hope that the bank is keeping a good record of the gains and losses of money in your account. How do you know if they are? One way is by checking the monthly balance statement sent to you by the bank. You can also compare the debits and credits in your passbook to the statement from the bank.

Just as your bank balance is important, a balanced equation for a redox reaction is important, too. You learned in the previous activity that oxidation numbers allow you to track the loss and gain of electrons in a reaction. What you need to do now is look at how oxidation numbers also allow you to balance the loss and gain of electrons in a redox reaction.

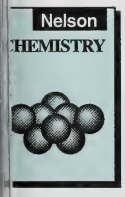
In Section 1 of this module you learned the five-step method of predicting and balancing redox reaction equations by using a redox table.

1. a. As a review of the redox table method, predict the most likely redox reaction from the following information:

   The orange-brown colour of liquid bromine disappears when mixed with hydrotelluric acid.

   b. What other evidence, besides a colour change, would support this as a predicted spontaneous reaction?

Check your answers by turning to the Appendix, Section 2: Activity 3.

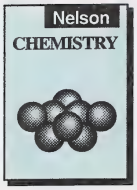


Why would you need another method of balancing redox reaction equations? To answer this question, read the top of page 376 in your textbook.

2. Why doesn't the redox table enable you to balance all redox equations?

3. What is the important assumption upon which both the half-reaction and oxidation number methods of balancing redox equations are based?

Check your answers by turning to the Appendix, Section 2: Activity 3.



As with the method of predicting redox equations you have already studied, your textbook outlines a five-step method for balancing redox equations using oxidation numbers. Carefully study the detailed steps on pages 377 and 378 of your textbook, including the examples given and the summary information. Once you have looked at the textbook examples, the following example will further clarify the balancing method.

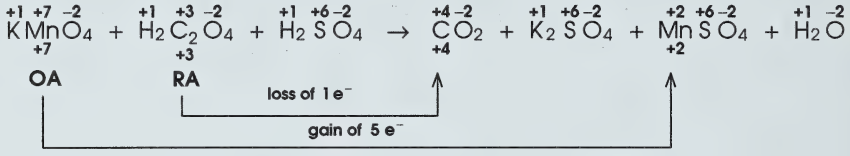
Example

Balance the following equation using oxidation numbers. (The states are omitted from the equation to simplify the work.)

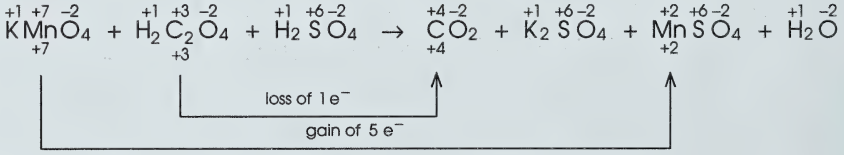
$$KMnO_4 + H_2C_2O_4 + H_2SO_4 \rightarrow CO_2 + K_2SO_4 + MnSO_4 + H_2O$$

Note: You may wish to try this example on your own first before reading further.

STEP 1 Assign oxidation numbers and note changes.

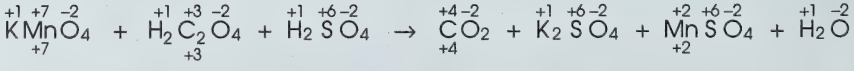


STEP 2 Determine the number of electrons transferred per atom or ion.



5 e⁻/Mn **1 e⁻/C**

STEP 3 Determine the number of electrons transferred per mole of OA and RA.



5 e⁻/Mn 1 e⁻/C

5 e⁻/$KMnO_4$ **2 e⁻/$H_2C_2O_4$**

STEP 4 Balance the number of electrons transferred by using appropriate coefficients for the OA and RA in the equation.

$$\mathbf{2}\ \overset{+1\ +7\ -2}{K\underset{+7}{Mn}O_4} + \mathbf{5}\ \overset{+1\ +3\ -2}{H_2\underset{+3}{C_2}O_4} + \overset{+1\ +6\ -2}{H_2SO_4} \rightarrow \overset{+4\ -2}{\underset{+4}{C}O_2} + \overset{+1\ +6\ -2}{K_2SO_4} + \overset{+2\ +6\ -2}{\underset{+2}{Mn}SO_4} + \overset{+1\ -2}{H_2O}$$

$5\ e^-/Mn$ $\quad 1\ e^-/C$

$5\ e^-/KMnO_4$ $\quad 2\ e^-/H_2C_2O_4$

$(\times\mathbf{2})$ $\quad (\times\mathbf{5})$

Note: When the electrons transferred are balanced, the equation shows that $2 \times 5 = \mathbf{10}$ electrons are transferred from RA to OA.

STEP 5 Balance the remaining species by inspection and by charge, if necessary.

$$\mathbf{2}\ KMnO_4 + \mathbf{5}\ H_2C_2O_4 + \mathbf{3}\ H_2SO_4 \rightarrow \mathbf{10}\ CO_2 + K_2SO_4 + \mathbf{2}\ MnSO_4 + \mathbf{8}\ H_2O$$

You have to be careful in the last step (Step 5) of the balancing because many things are being balanced. Take special note again of the two paragraphs under Step 5 on page 378 of your textbook.

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4. Practise the oxidation number balancing method by doing textbook question 46 (a-h) on page 379 of *Nelson Chemistry*.

Check your answers by turning to the Appendix, Section 2: Activity 3.

For each of the equations that you balanced in the preceding question, do you know how many total electrons were transferred? The number of electrons transferred in a reaction will become important to you when dealing with the energy associated with redox reactions. So what is the answer? Study the following example.

net ionic equation – a balanced chemical equation containing no spectator atoms or ions

Example

You are given the following **net ionic equation**:

$$3\ Cu_{(s)} + 6\ H^+_{(aq)} + 2\ HNO_{3(aq)} \rightarrow 3\ Cu^{2+}_{(aq)} + 2\ NO_{(g)} + 4\ H_2O_{(l)},$$

a. What is the total number of electrons transferred?

b. How many moles of electrons are transferred if 2.3 mol of the reducing agent reacts?

To answer these questions, consider the number of electrons involved by assigning oxidation numbers and indicating the changes for the oxidizing and reducing agents.

RA OA

$$\overset{0}{3\ Cu}_{(s)} + 6\ \overset{+1}{H}^+_{(aq)} + 2\ \overset{+1+5\ -2}{HNO}_{3(aq)} \rightarrow 3\ \overset{+2}{Cu}^{2+}_{(aq)} + 2\ \overset{+2\ -2}{NO}_{(g)} + 4\ \overset{+1\ -2}{H_2O}_{(l)}$$

$0 \rightarrow +2$ (Cu) $\quad +5 \rightarrow +2$ (N)

$2\ e^-/Cu$ (×3) ↑ means that 2 mol e^- are transferred per mole of Cu reacted

$3\ e^-/N$

$3\ e^-/HNO_3$ (×2) ← means that in the balanced equation, $3 \times 2 = 6$ mol e^- are transferred in total

a. The total number of electrons transferred is 6 mol.

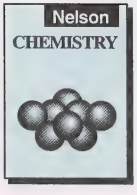
$$2\ e^-/Cu \times 3 = 6 \quad \textbf{or} \quad 3\ e^-/HNO_3 \times 2 = 6$$

b. For 2.3 mol of reducing agent $(Cu_{(s)})$ 4.6 mol e^- are transferred.

There are 2 mol e^- transferred per one mole of Cu. Therefore,

$$\frac{2\ mol\ e^-}{1\ \cancel{mol\ Cu}} \times 2.3\ \cancel{mol\ Cu}$$

$= 4.6$ mol e^- transferred



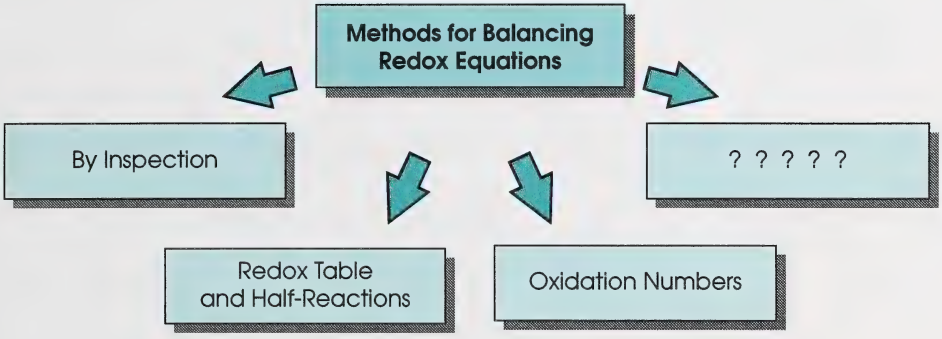
5. For each of the redox equations you balanced in textbook question 46 (a-h) on page 379 of your textbook, indicate the total number of electrons transferred.

6. Use the following net ionic equation to answer this question.

$$2\ Cr_2O_7^{2-}{}_{(aq)} + 16\ H^+{}_{(aq)} + 3\ C_2H_5OH_{(aq)} \rightarrow 4\ Cr^{3+}{}_{(aq)} + 11\ H_2O_{(l)} + 3\ CH_3COOH_{(aq)}$$

If 35.2 g of reducing agent reacts, how many moles of electrons are transferred in the reaction?

Check your answers by turning to the Appendix, Section 2: Activity 3.



Every so often, you will run into a redox equation which just doesn't allow you to balance the equation by one of the three methods you have studied so far. When this happens, you have one more method to fall back on – balancing the equation by devising your *own* half-reactions.

An example of an equation where you might need to devise your own half-reactions is the following:

$$S_2O_3^{2-}{}_{(aq)} + OCl^-{}_{(aq)} \rightarrow Cl^-{}_{(aq)} + S_4O_6^{2-}{}_{(aq)} \text{ (acidic)}$$

This is a case where the following conditions exist:

- The half-reactions are not listed in the redox table.
- You are told the reaction occurs in acidic solution.
- You don't have all the species given that actually are in the overall equation.

The following method will enable you to balance this type of redox equation. A question using this example might be worded as shown.

Example

Balance the following equation. The reaction occurs in an acidic solution.

$$S_2O_3^{2-}(aq) + OCl^-(aq) \rightarrow Cl^-(aq) + S_4O_6^{2-}(aq)$$

STEP 1 Separate the equation into two half-reactions. (You can identify which elements are involved by using oxidation numbers, if you wish.)

$$\left(\overset{+2}{S_2}\overset{-2}{O_3}^{2-}(aq) + \overset{-2}{O}\overset{+1}{Cl}^-(aq) \rightarrow \overset{-1}{Cl}^-(aq) + \overset{+\frac{5}{2}}{S_4}\overset{-2}{O_6}^{2-}(aq) \right)$$

$$\mathbf{S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq)}$$

$$\mathbf{OCl^-(aq) \rightarrow Cl^-(aq)}$$

STEP 2 Balance the species, other than oxygen and hydrogen, in each half-reaction.

$$\mathbf{2}\ S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq)$$

$$OCl^-(aq) \rightarrow Cl^-(aq)$$

STEP 3 Balance oxygen atoms by adding sufficient $H_2O_{(l)}$ to the appropriate side of the equations. (This is justified since the reaction is occurring in aqueous solution.)

$$2\ S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq) \quad \leftarrow \textbf{O's balanced already}$$

$$OCl^-(aq) \rightarrow Cl^-(aq) + \mathbf{H_2O_{(l)}}$$

STEP 4 Balance hydrogen atoms by adding sufficient $H^+(aq)$ ions to the appropriate side of the equations. (This is justified since acidic conditions are present.)

$$2\ S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq) \quad \leftarrow \textbf{no H's present}$$

$$\mathbf{2\ H^+(aq)} + OCl^-(aq) \rightarrow Cl^-(aq) + H_2O_{(l)}$$

STEP 5 Add sufficient electrons to the appropriate side of each equation to balance charge.

$$2\,S_2O_3^{2-}{}_{(aq)} \rightarrow S_4O_6^{2-}{}_{(aq)} + \mathbf{2\,e^-}$$
$$2(2^-) = \mathbf{4^-} \qquad 2^- + 2^- = \mathbf{4^-}$$

$$\mathbf{2\,e^-} + 2\,H^+{}_{(aq)} + OCl^-{}_{(aq)} \rightarrow Cl^-{}_{(aq)} + H_2O_{(l)}$$
$$2^- + 2(1^+) + 1^- = \mathbf{1^-} \qquad 1^- + 0 = \mathbf{1^-}$$

STEP 6 Multiply each half-reaction by the appropriate factor to balance electrons transferred (just like the half-reaction method you studied earlier).

$$2\,S_2O_3^{2-}{}_{(aq)} \rightarrow S_4O_6^{2-}{}_{(aq)} + \mathbf{2\,e^-}$$
$$\mathbf{2\,e^-} + 2\,H^+{}_{(aq)} + OCl^-{}_{(aq)} \rightarrow Cl^-{}_{(aq)} + H_2O_{(l)}$$

Electrons transferred are already balanced.

STEP 7 Add the balanced half-reaction equations to get an overall equation.

$$2\,S_2O_3^{2-}{}_{(aq)} \rightarrow S_4O_6^{2-}{}_{(aq)} + \cancel{2e^-}$$
$$\cancel{2e^-} + 2\,H^+{}_{(aq)} + OCl^-{}_{(aq)} \rightarrow Cl^-{}_{(aq)} + H_2O_{(l)}$$
$$\mathbf{2\,S_2O_3^{2-}{}_{(aq)} + 2\,H^+{}_{(aq)} + OCl^-{}_{(aq)} \rightarrow S_4O_6^{2-}{}_{(aq)} + Cl^-{}_{(aq)} + H_2O_{(l)}}$$

To give you some practice with this balancing method, answer the following question.

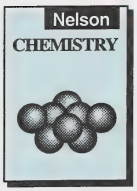
7. Balance the following redox equations by devising your own half-reactions. Each reaction occurs in an acidic, aqueous solution.

   a. $ClO_3^-{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow Cl_{2(g)} + ClO_{2(aq)}$

   b. $Cu_{(s)} + HNO_{3(aq)} \rightarrow Cu^{2+}{}_{(aq)} + NO_{(g)}$

   c. $Fe^{2+}{}_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow Fe^{3+}{}_{(aq)} + Mn^{2+}{}_{(aq)}$

Check your answers by turning to the Appendix, Section 2: Activity 3.



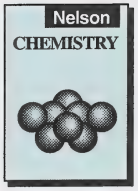
One of the many practical applications of the redox concepts you have been studying is the analysis of blood alcohol content by means of a redox reaction. The same redox reaction is used by police in roadside breathalysers. A redox reaction of alcohol in the blood or of exhaled air with potassium dichromate allows quick detection of alcohol levels. To gain further insight into these applications, first read the description of breathalysers on page 380 of your textbook.

8. When driving, what is the legal limit for blood alcohol content in Canada?

9. Why are breathalysers used?

10. Briefly describe how a breathalyser works.

Note: The following is a hypothetical situation.



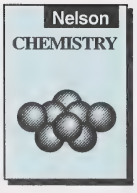
Suppose that you are doing the analysis of a blood sample taken from a suspected impaired driver. Instead of using the standard procedure for blood alcohol testing (as is described in Lab Exercise 12J: Analysing Blood for Alcohol Content on pages 379 and 380 of your textbook), you have found a more direct approach. You directly titrate the alcohol (ethanol) in a blood sample with acidified potassium dichromate solution. The partial equation for the dichromate/ethanol reaction that occurs is as follows:

$$\underset{\left(\begin{matrix}\text{dichromate}\\ \text{ion}\end{matrix}\right)}{Cr_2O_7^{2-}{}_{(aq)}} + \underset{\text{(ethanol)}}{C_2H_5OH_{(aq)}} \rightarrow Cr^{3+}{}_{(aq)} + CH_3COOH_{(aq)} \quad \text{(unbalanced)}$$

Remember, the reaction is carried out in acidic conditions.

11. Balance the preceding equation by devising your own half-reactions.

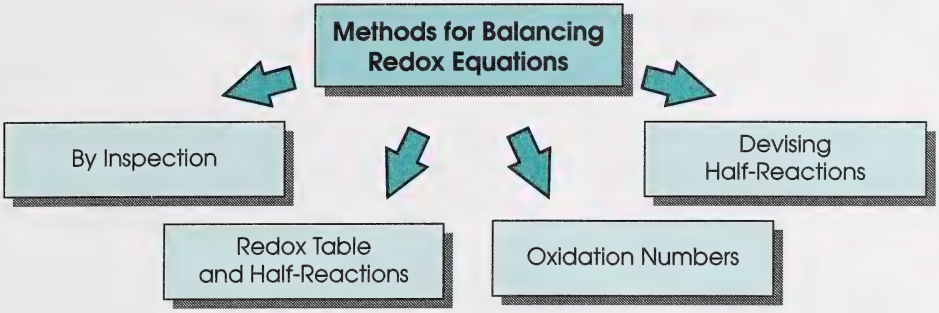
12. If 3.66 mL of 2.15 mmol/L acidified potassium dichromate was required to completely react with the ethanol in a 0.500 mL blood sample, what was the molar concentration of the $C_2H_5OH_{(aq)}$?



13. a. Use Table 12.5 on page 380 of your textbook to construct a graph of molar concentration of ethanol versus blood alcohol content (manipulated variable). Then, from the graph, find the blood alcohol content (mg/100 mL) equivalent to the molar concentration calculated in the preceding question.

    b. Based on your findings, was the driver legally impaired? Explain.

Check your answers by turning to the Appendix, Section 2: Activity 3.



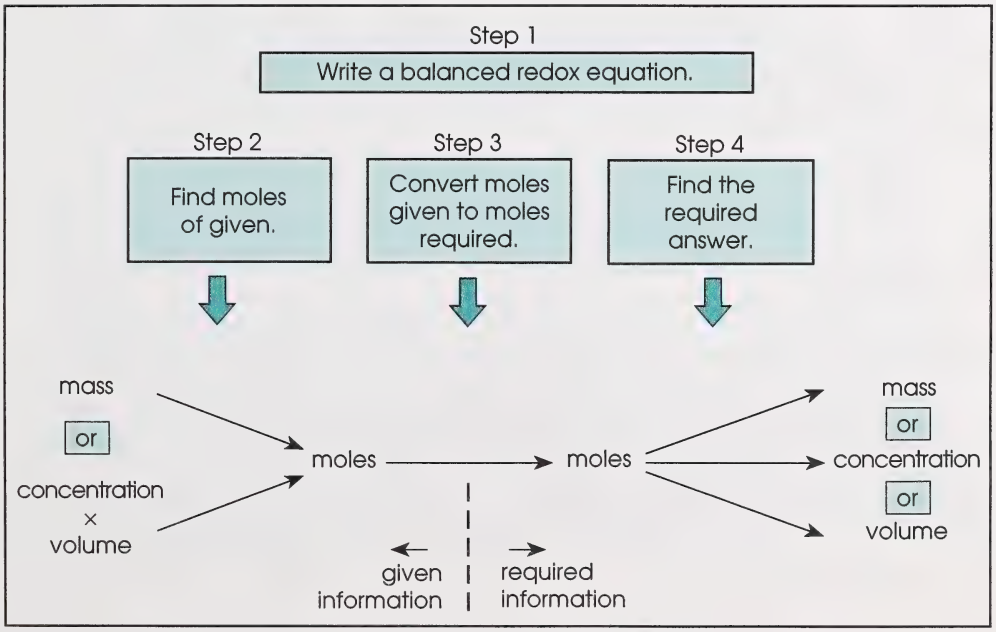
The methods of balancing equations that you have studied will provide you with the means of predicting and balancing most redox reaction equations.

Follow-up Activities

If you had difficulties understanding the concepts in the activities, it is recommended that you do the Extra Help. If you have a clear understanding of the concepts, it is recommended that you do the Enrichment.

Extra Help

The steps used to solve redox stoichiometry problems are as follows:



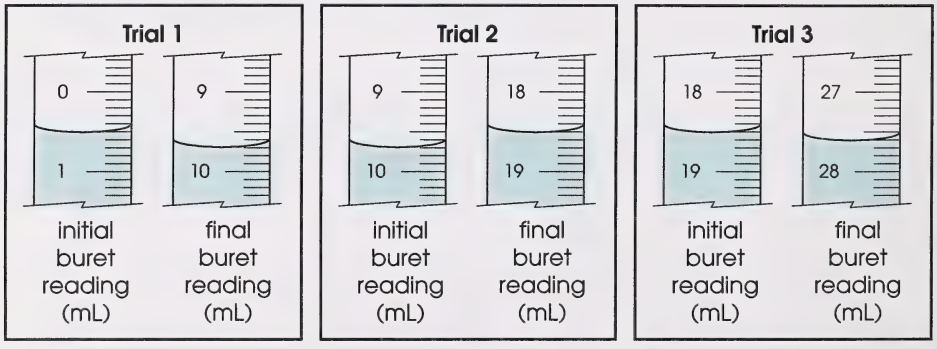
1. Match each of the following terms with its correct defining statement.

a. titration	i. the point in a titration at which chemically equivalent amounts of titrant and sample being titrated are present
b. titrant	ii. slow and precise addition of a solution in a buret to a measured amount of a sample
c. equivalence point	iii. the buret solution used in a titration
d. endpoint	iv. the point in a titration when an observable change takes place, e.g., colour change
e. primary standard	v. a pure and stable form of a chemical used to find the concentration of a reactant solution

2. What colour change would you expect the $MnO_4^-{}_{(aq)}$ ion to undergo when it is reduced in a redox reaction?

3. Another oxidizing agent other than the $MnO_4^-{}_{(aq)}$ ion which undergoes a colour change from orange to green when it is reduced would be _____.

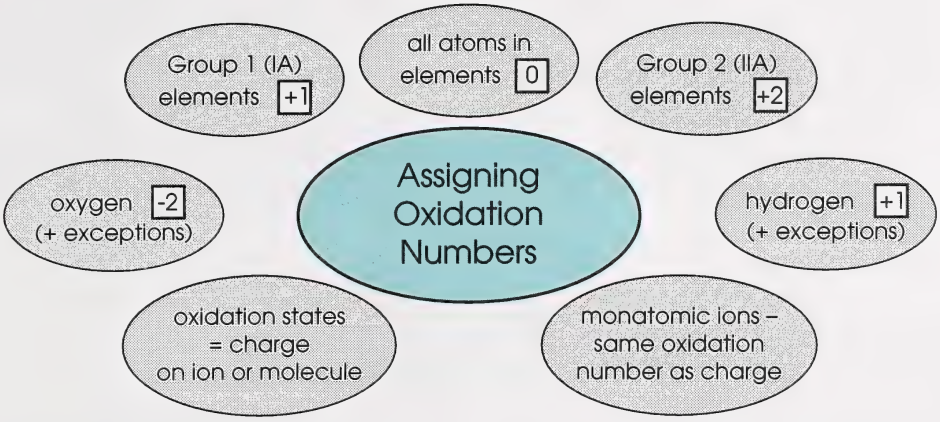
4. Using the following data from a titration of $Sn^{2+}{}_{(aq)}$ with acidified $KMnO_{4(aq)}$, calculate the average volume of acidified $KMnO_{4(aq)}$ solution that was used.



Endpoint colour – pink | Endpoint colour – pink | Endpoint colour – pink

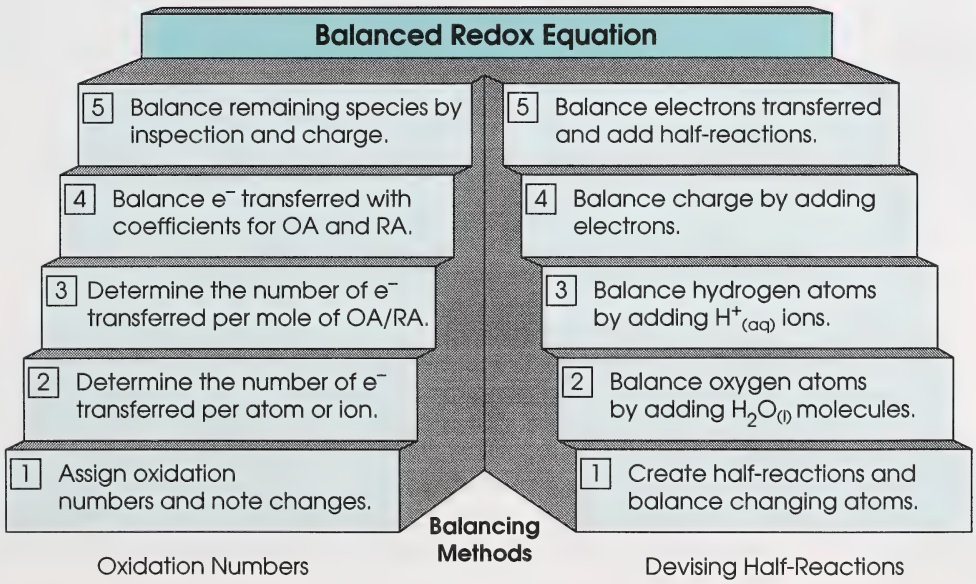
5. If 10.0 mL of 0.555 mol/L of $SnCl_{2(aq)}$ were reacted with the calculated volume of $KMnO_{4(aq)}$ from question 4, what was the concentration of $KMnO_{4(aq)}$?

The following graphic illustrates the basic rules for assigning oxidation numbers to elements.


Assigning Oxidation Numbers
Group 1 (IA) elements +1
all atoms in elements 0
Group 2 (IIA) elements +2
oxygen -2 (+ exceptions)
hydrogen +1 (+ exceptions)
oxidation states = charge on ion or molecule
monatomic ions – same oxidation number as charge

6. List the following species in order of increasing oxidation number of the sulphur atom: H_2S, S_8, H_2SO_4, S^{2-}, SO_2

The following graphic illustrates the basic steps involved in balancing redox equations by assigning oxidation numbers and by devising half-reactions.


Balanced Redox Equation
5 Balance remaining species by inspection and charge.
4 Balance e⁻ transferred with coefficients for OA and RA.
3 Determine the number of e⁻ transferred per mole of OA/RA.
2 Determine the number of e⁻ transferred per atom or ion.
1 Assign oxidation numbers and note changes.
5 Balance electrons transferred and add half-reactions.
4 Balance charge by adding electrons.
3 Balance hydrogen atoms by adding H+(aq) ions.
2 Balance oxygen atoms by adding H2O(l) molecules.
1 Create half-reactions and balance changing atoms.
Balancing Methods
Oxidation Numbers
Devising Half-Reactions

7. Balance the following equations, using one of the redox methods you learned in this module.

   a. $Cr^{3+}_{(aq)} + BiO_3^-{}_{(aq)} \rightarrow Cr_2O_7^{2-}{}_{(aq)} + Bi^{3+}_{(aq)}$ (in acidic solution)

   b. $Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_2S_{(aq)} + H_2O_{(l)}$

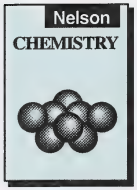
8. For each of the following complete redox reactions, write the reduction and oxidation half-reactions and identify the oxidizing agent and the reducing agent.

   a. $2\ Sr_{(s)} + O_{2(g)} \rightarrow 2\ SrO_{(s)}$

   b. $2\ Na_{(s)} + 2\ H_2O_{(l)} \rightarrow H_{2(g)} + 2\ NaOH_{(aq)}$

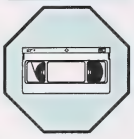
   c. $Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

Check your answers by turning to the Appendix, Section 2: Extra Help.



If you would like further practice in working with redox questions, you can do any of textbook questions 7, 8, 9, 17, 18, 19, 22, 23, 24, and 25 on pages 382-384 of *Nelson Chemistry*.

The answers to these questions can be found on pages 515-516 in your textbook.



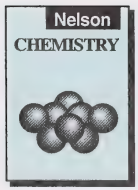
If you have access to the video *Redox* (Classroom Video), watch the first twenty minutes to review the concepts studied in this module.

Enrichment

Do **one** of the following activities.

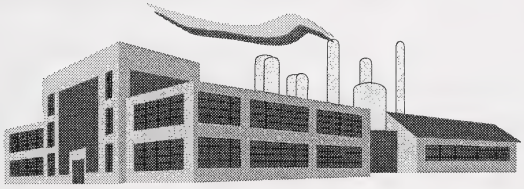
1. Nitric acid is a strong oxidizing agent. Explain which of the following would be least likely to be produced from a reaction of nitric acid and zinc metal:

   $$NO,\ NO_2,\ N_2O_4,\ N_2O_5,\ NH_4^+$$



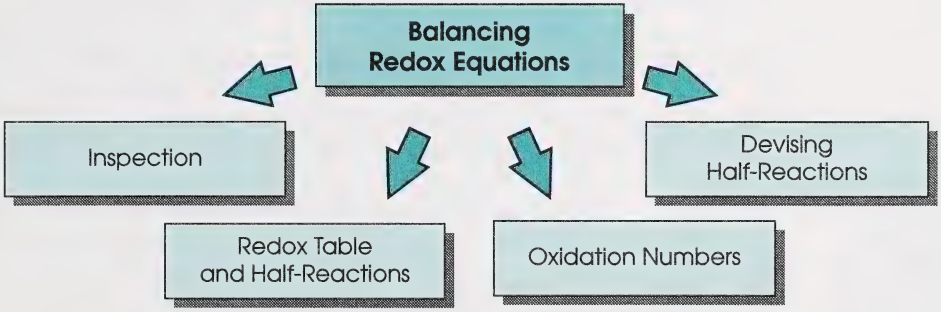
2. Complete the analysis of Lab Exercise 12K: Analysing Antifreeze on page 385 of your textbook.

3. In the production of steel, small amounts of carbon and manganese are mixed with iron. To determine the amount of manganese present, a sample of steel is initially dissolved in concentrated nitric acid; the reaction produces $Mn^{2+}_{(aq)}$ and $NO_{2(g)}$. Next, the $Mn^{2+}_{(aq)}$ ion is reacted with an acidic solution containing $IO_4^-{}_{(aq)}$ to give $MnO_4^-{}_{(aq)}$ and $IO_3^-{}_{(aq)}$. Finally the concentration of the permanganate solution is determined by titrating it with a standard acidified $Fe^{2+}_{(aq)}$ solution to produce $Mn^{2+}_{(aq)}$ and $Fe^{3+}_{(aq)}$. Write balanced equations for each step.



Check your answers by turning to the Appendix, Section 2: Enrichment.

Conclusion



Redox reactions can be balanced using a number of different methods. Your decision on which method to use is to be largely based on whether or not you need to predict products. Oxidation numbers, which are assigned values to chemical species, are very useful in determining whether species are gaining or losing electrons. They also allow you to quickly determine whether a reaction is of the redox type or not. Lastly, they allow you to balance certain kinds of redox reactions. You have also seen the application of stoichiometry to redox reactions in determining amounts of reactant material and in the analysis of reagents.

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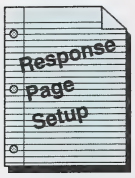
Section 2 Assignment: Redox Applications

Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

Chemistry 30 – Module 3 Section 2 Assignment Page # Name and ID#

Be sure to write legibly. Leave a wide left margin and number all of your pages.



1. An increase in the oxidation number of a chemical species causes it to be _____ (oxidized/reduced). Write the single-word answer on your answer sheet. **(1 mark)**

2. Explain if it is possible to have a reaction in which oxidation takes place, but reduction does not. **(2 marks)**

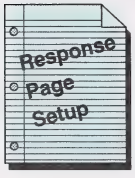
3. A student titrated a sample of an acidified tin(II) chloride solution with potassium permanganate.

   a. List the oxidizing agent and reducing agent that would participate in the reaction. **(2 marks)**

   b. What colour change would signify the endpoint? **(1 mark)**

   c. Which chemical species is the titrant? **(1 mark)**

   d. When would the equivalence point of this titration be reached? Be specific. Give the actual mole ratio of the oxidizing and reducing agents. **(2 marks)**



4. Indicate for each of the following equations whether or not they represent a redox reaction. Write *redox* or *not redox* as your answer choice for each equation. **(4 marks)**

   a. $Ca_{(s)} + F_{2(g)} \rightarrow CaF_{2(s)}$

   b. $MgCl_{2(aq)} + 2\ NaOH_{(aq)} \rightarrow Mg(OH)_{2(s)} + 2\ NaCl_{(aq)}$

   c. $CaCO_{3(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

   d. $NaI_{(aq)} + 3\ HOCl_{(aq)} \rightarrow NaIO_{3(aq)} + 3\ HCl_{(aq)}$

5. Indicate the oxidation number of phosphorous in each of the following species. For each case, write P = and then your answer. **(4 marks)**

   a. HPO_3^{2-} b. H_3PO_2 c. P_4O_{10} d. Na_2HPO_4

Response Page Setup

6. Write the balanced chemical equation for the redox reaction that might occur when hydrogen gas is bubbled into a basic solution of gold(III) nitrate.

   In your response, show the half-reactions, the overall equation, and predict the spontaneity of the reaction. **(3 marks)**

7. Balance the following redox equation by using the oxidation number method. In your response, indicate the appropriate oxidation numbers and balancing coefficients. **(3 marks)**

$$NaIO_{3(aq)} + NaI_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + I_{2(s)} + H_2O_{(l)}$$

Response Page Setup

8. Balance the following redox equation in acidic solution by devising half-reactions. **(3 marks)**

$$H_3AsO_{3(aq)} + Cr_2O_7^{2-}{}_{(aq)} \rightarrow H_3AsO_{4(aq)} + Cr^{3+}{}_{(aq)}$$

   In your response, include half-reaction equations and the overall balanced equation.

Response Page Setup

9. This question is a numerical response question. Record only your final answer to three digits in a box similar to the following:

   Enter the first digit of your answer in the left-hand box and leave any unused boxes blank. If a decimal point occurs in your answer, it should occupy its own box.

   When 1.20 mol of $H_2PO_4^-{}_{(aq)}$ reacts according to the following equation

$$12\ H_2PO_4^-{}_{(aq)} + 12\ H^+{}_{(aq)} + 10\ Br^-{}_{(aq)} \rightarrow 3\ P_{4(s)} + 18\ H_2O_{(l)} + 10\ BrO_3^-{}_{(aq)}$$

   the number of moles of electrons transferred is _____ mol. **(2 marks)**

10. Use the following information to answer this question.

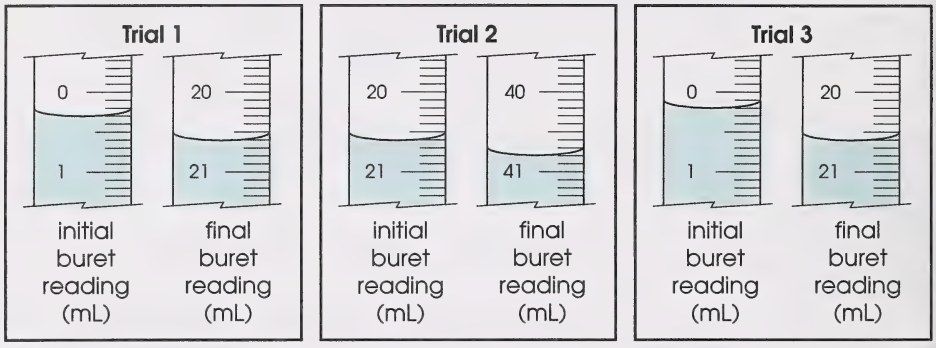
$$PCl_{5(s)} + 2\,Na_{(s)} \rightarrow PCl_{3(s)} + 2\,NaCl_{(s)} + 729.5\ kJ$$

Calculate the heat produced when 0.235 mol of reducing reagent reacts. Show all your work. **(2 marks)**

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☐ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☐ F. Evaluating

11. A chemistry student titrated three 25.0 mL samples of acidified tin(II) nitrate solution with 0.00800 mol/L potassium permanganate to find the concentration of the $Sn^{2+}_{(aq)}$ ions in the solution. Obtain the results of the three trials by reading the buret volumes (in millilitres) of $KMnO_{4(aq)}$.



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Response Page Setup

Based on the data given, calculate the concentration of $Sn^{2+}_{(aq)}$ ions in the solution. Show all your work in your response, including a net redox equation and calculations. **(6 marks)**

Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

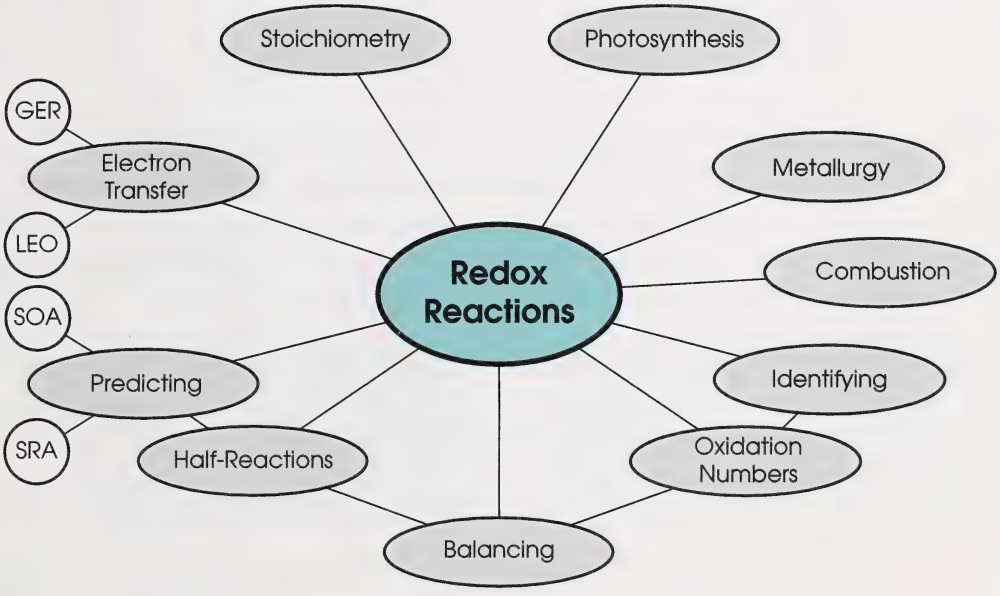
Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

Response Page Setup

12. Iron metal is easily oxidized to $Fe^{2+}_{(aq)}$ by an acidic potassium dichromate solution. Write the net equation for this process and determine the mass of iron metal oxidized by 50.0 mL of 0.250 mol/L acidified $K_2Cr_2O_{7(aq)}$. Show all your work. **(4 marks)**

MODULE SUMMARY

You began this module by reading about the downfall of a king due to the progress of technology. The technology that you dealt with in this module is associated with electrochemistry. Redox or electrochemical reactions transfer electrons. You learned how these transfers occur and what terms chemists use to describe them. You discovered that many reaction types that you originally learned about in Chemistry 20 can actually be classified as redox reactions. Science makes great use of these reactions to provide you with many of the things you take for granted on a day-to-day basis. Since the industrial and scientific applications of redox principles are widespread, the concepts you have studied provide you with a means to better understand the world around you.



20

Final Module Assignment

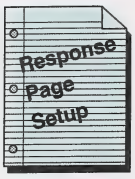
Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

Chemistry 30 – Module 3 Final Module Assignment Page # Name and ID#

Be sure to write legibly. Leave a wide left margin and number all of your pages.

Part A: Multiple Choice



Read each question carefully and decide which of the choices **best** completes the statement or answers the question. Place the number of each question in a vertical column and place your answer (A, B, C, or D) beside each number. Each question is worth one mark. **(Total = 5 marks)**

1. Students tried the following combination of reactants in a laboratory. Which combination is expected to be non-spontaneous?

   A. $Ag^{+}_{(aq)} + H_{2(g)}$

   B. $Cr_{(s)} + Sn^{2+}_{(aq)}$

   C. $Ba_{(s)} + Fe^{2+}_{(aq)}$

   D. $Al_{(s)} + Ca^{2+}_{(aq)}$

2. Jody wishes to prepare 100 mL of a 0.010 mol/L $SnF_{2(aq)}$ to titrate with freshly acidified $KMnO_{4(aq)}$ of uncertain concentration. The equipment available to Jody includes a 100 mL volumetric flask, a 10 mL pipette, a balance with a precision of ±0.01 g, and a 150 mL beaker. Once Jody obtains the appropriate mass of $SnF_{2(s)}$, which procedure should she use to prepare this solution?

   A. Place the $SnF_{2(s)}$ in the beaker and add exactly 100 mL of water from the volumetric flask.
   B. Place the $SnF_{2(s)}$ in the beaker and add exactly 100 mL of water from the pipette in 10 mL portions.
   C. Place the $SnF_{2(s)}$ in the volumetric flask, dissolve it in less than 100 mL of water, and then dilute it to the 100 mL mark.
   D. Place the $SnF_{2(s)}$ in the beaker, dissolve it in more than 100 mL of water, and then pour the solution into the volumetric flask to the 100 mL mark.

3. The following equation shows the burning of hydrogen cyanide.

$$4\ HCN_{(g)} + 9\ O_{2(g)} \rightarrow 4\ CO_{2(g)} + 2\ H_2O_{(g)} + 4\ NO_{2(g)}$$

   Which atom does **not** change its oxidation state?

   A. hydrogen
   B. carbon
   C. oxygen
   D. nitrogen

4. You are given the following spontaneous reactions:

$$X^{2+}_{(aq)} + Y_{(s)} \rightarrow X_{(s)} + Y^{2+}_{(aq)}$$
$$X^{2+}_{(aq)} + Q_{(s)} \rightarrow X_{(s)} + Q^{2+}_{(aq)}$$
$$Y^{2+}_{(aq)} + Z_{(s)} \rightarrow Y_{(s)} + Z^{2+}_{(aq)}$$
$$Z^{2+}_{(aq)} + Q_{(s)} \rightarrow Z_{(s)} + Q^{2+}_{(aq)}$$

The strongest oxidizing agent is

A. $Q^{2+}_{(aq)}$
B. $Y^{2+}_{(aq)}$
C. $Z^{2+}_{(aq)}$
D. $X^{2+}_{(aq)}$

5. Use the following information to answer this question.

A student made these observations after mixing several test solutions with samples of TTE (1,1,2-trichloro-1,2,2-trifluoroethane), a hydrocarbon solvent.

Solution	Colour of TTE Layer
$Cl^-_{(aq)}$	colourless
$Br^-_{(aq)}$	colourless
$I^-_{(aq)}$	colourless

Solution	Colour of TTE Layer
$Cl_{2(aq)}$	colourless
$Br_{2(aq)}$	orange
$I_{2(aq)}$	pink

She then mixed various combinations of solutions and obtained these results:

I. $Cl_{2(aq)} + Br^-_{(aq)}$: TTE layer orange

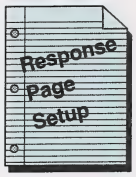
II. $Cl^-_{(aq)} + Br_{2(aq)}$: TTE layer orange

III. $I^-_{(aq)} + Br_{2(aq)}$: TTE layer pink

The reaction(s) that shows that bromine was produced is(are)

A. I only
B. III only
C. I and II
D. I, II, and III

Part B: Numerical Response



Read each question carefully. Record only your final answer by writing it in a four-box unit similar to the following: ☐☐☐☐

Enter the first digit of your answer in the left-hand box and leave any unused boxes blank. If a decimal point occurs in your answer, it should occupy its own box.

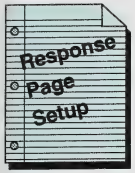
6. The following hypothetical metals were placed in solutions of ions of the other three metals. Mixtures that showed evidence of a chemical reaction were indicated by a check mark.

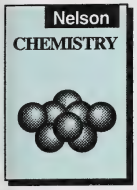
Metals \ Metal ions		1 $E^{+}_{(aq)}$	2 $A^{2+}_{(aq)}$	3 $D^{2+}_{(aq)}$	4 $J^{3+}_{(aq)}$
1	$E_{(s)}$	_____	_____	✔	_____
2	$A_{(s)}$	✔	_____	✔	✔
3	$D_{(s)}$	_____	_____	_____	_____
4	$J_{(s)}$	✔	_____	✔	_____

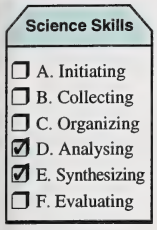
The reducing agents, in decreasing order of strength, are _____. **(Record all four agents by their numbers.) (1 mark)**

7. The oxidation number of the sulphur in each of the compounds Na_2SO_4, Na_2SO_3, SO_2, and SO_3 respectively is _____. (Record all four digits in the boxes on your response sheet.) **(1 mark)**

Part C: Written Response







Write your answers neatly and concisely. For full marks, your answers must show all pertinent explanations, calculations, and formulas.

8. A simple method of removing the dark tarnish, $Ag_2S_{(s)}$, from silver objects was explained on page 377 of your textbook. Would the method work if the tarnished objects were placed in a copper pan instead of an aluminum pan?

   Justify your answer. In your response, use all pertinent reaction equations and explanations. **(4 marks)**

9. An investigation was performed by a student in which 0.45 g of aluminum was to be completely reacted in a copper(II) chloride solution. The purpose of the lab was to determine the relationship between the amount of heat released by the reaction and the moles of electrons involved in the reaction as it released its heat.

   Based on the following procedure and evidence, you will calculate the heat released and the relationship between the heat released and electrons transferred, as required in the analysis questions that follow.

Investigation: An Exothermic Redox Reaction

Problem

What quantity of heat energy is released when aluminum foil reacts with a solution of copper(II) chloride?

Procedure

- Obtain 0.45 g of aluminum and crumple the aluminum into a loose ball.
- Calculate the mass of copper(II) chloride required to prepare 100.0 mL of a 0.500 mol/L solution.
- Obtain the required mass of copper(II) chloride.
- Dissolve the required amount of copper(II) chloride in 100.0 mL of distilled water in a polystyrene cup calorimeter.
- Take the initial temperature of the $CuCl_{2(aq)}$ solution after it has dissolved.
- Add the aluminum to the solution.
- Take the final temperature of the solution after reaction has been completed.

Evidence

mass of aluminum reacted = 0.45 g
volume of solution = 100.0 mL
initial temperature of solution = 24.0°C
final temperature of solution = 46.5°C

Assumptions:

- No heat is lost to the surroundings.
- The solute occupies no volume.
- The solution has the same specific heat capacity as water.

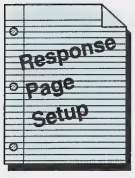
Analysis

a. Write the net ionic redox equation for the reaction. **(2 marks)**

b. Calculate the molar heat of reaction of aluminum. **(4 marks)**

c. Write the net balanced thermochemical equation. **(1 mark)**

d. Calculate the heat released per mole of electrons transferred. **(2 marks)**



Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Leave a space for the teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

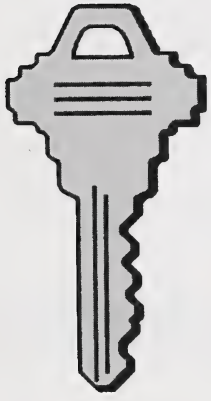
To ensure that all of your work has been completed in a satisfactory manner, check off the items in the following list:

- ❑ Section 1 Assignment has been completed.
- ❑ Section 2 Assignment has been completed.
- ❑ Final Module Assignment has been completed.
- ❑ Your responses are organized and neat, with room for teacher comments.
- ❑ All your response pages are numbered consecutively and identified with this heading.

Chemistry 30 – Module 3 Section # Assignment Page # Name and ID #

Submit **only** your **assignment response pages** for evaluation.

Appendix



Glossary
Suggested Answers

Glossary

endpoint: the point in a titration at which an observable change occurs, e.g., colour

equivalence point: the point in titration at which chemically equivalent amounts of titrant and sample are present

half-reaction equations: balanced equations that describe one-half of an overall process

half-reactions in acid solution: method used to balance redox reactions that occur in acidic solution

mnemonic: a memory aid

net ionic equation: a balanced chemical equation containing no spectator atoms or ions

oxidation: the process by which electrons are lost in a reaction

oxidation number: a positive or negative number representing the oxidation state of an atom

oxidizing agent: a chemical species that accepts electrons in a reaction

primary standard: a stable chemical that can be used in a standard solution

reduction: the process by which electrons are gained in a reaction

reducing agent: a chemical species that donates electrons in a reaction

spontaneous: occurring naturally without outside help

titrant: the buret solution used in a titration

titration: the accurate addition of a solution in a buret to a measured amount of a sample

Suggested Answers

Section 1: Activity 1

1. It is most likely that metals such as gold and silver were found in pure form in river beds and other sites exposed by erosion.

2. Iron would have been used in the weapons of the invading army.

3. a. The invaders' advanced metallurgical technology allowed them to take large amounts of iron ore and produce smaller amounts of iron by a process called **reduction**.

   b. Other common metals produced by reduction would have been copper and tin.

4. To produce the metals that were used to make weapons, a substance had to be used to help the reaction occur. This substance would generally be called a **reducing agent**.

5. Halogens such as chlorine, bromine, or iodine can oxidize metals.

6. **Textbook question 2:**

   a. The reaction shows the oxidation of iron; oxygen is the oxidizing agent.

b. The reaction shows reduction of lead(II) oxide; carbon is the reducing agent.

c. The reaction shows reduction of nickel(II) oxide; hydrogen is the reducing agent.

d. The reaction shows the oxidation of tin; bromine is the oxidizing agent.

e. The reaction shows reduction of iron(III) oxide; carbon monoxide is the reducing agent.

f. The reaction shows the oxidation of copper; nitric acid (aqueous hydrogen nitrate) is the oxidizing agent.

7. The predicted balanced equations are as follows:

$Cu_{(s)} + 2\ AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2\ Ag_{(s)}$

$Cl_{2(aq)} + 2\ NaBr_{(aq)} \rightarrow 2\ NaCl_{(aq)} + Br_{2(l)}$*

$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

$Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$

$Cl_{2(aq)} + 2\ KI_{(aq)} \rightarrow 2\ KCl_{(aq)} + I_{2(s)}$*

*Since both Br_2 and I_2 dissolve somewhat in water, their states could also be written as aqueous $_{(aq)}$.

8. A description of the design is as follows: The reactions are performed as outlined in the problem and evidence is collected to confirm the predictions made, including diagnostic tests.

The predictions will be confirmed as follows:

- Silver will be a silvery solid and the copper(II) nitrate will be a blue-coloured solution in Reaction 1.
- Bromine will be an orange colour in a solvent layer of a halogen diagnostic test in Reaction 2.
- Hydrogen will give a pop sound when ignited in a diagnostic test in Reaction 3.
- Copper will be a red-brown solid formed in Reaction 4.
- Iodine will be a purple colour in the solvent layer of a halogen diagnostic test in Reaction 5.

9. A possible evidence table is as follows:

Test Tube Number	Initial Observations after Reaction	Observations from Diagnostic Tests
1	• A silvery solid formed on copper. • There was a blue coloured solution.	• not applicable
2	• There was a yellow-brown coloured solution.	• The solvent layer was orange in colour.
3	• A gas bubbled off the magnesium. • The solution remained colourless.	• A pop sound was heard when the gas was ignited.
4	• A reddish brown solid formed on the zinc.	• not applicable
5	• There was a yellow-brown coloured solution.	• The solvent layer was violet (purple) in colour.

10. Refer to the answer for Section 1: Activity 1, question 9.

11.

Test Tube Number	Analysis – Confirmation of Products
1	$Ag_{(s)}$ was confirmed by a silvery solid. $Cu(NO_3)_{2(aq)}$ was confirmed by a blue solution.
2	$Br_{2(aq)}$ was confirmed by an orange solvent layer.
3	$H_{2(g)}$ was confirmed by the pop sound when the $H_{2(g)}$ was ignited.
4	$Cu_{(s)}$ was confirmed by a reddish brown solid on the zinc.
5	$I_{2(aq)}$ was confirmed by a purple solvent layer.

12. A possible evaluation is as follows:

The experimental design and procedure seemed adequate to answer the problem since at least one product was confirmed for each reaction studied. There were no conflicting results from the diagnostic tests, which supports the design. More confidence could be placed in the design, however, if every product was confirmed through further diagnostic testing. The procedure was adequate since it allowed for the collection of all data indicated in the design.

13. **Textbook question 6:**

a. Copper(II) ions are converted to copper atoms.

b. $Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$

c. Zinc atoms are converted to zinc ions.

d. $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^-$

e. Copper ions gain electrons and the zinc atoms lose electrons.

f. Sulphate ions do not change during this reaction. A substance (ion) that does not change or take part during a reaction is called a spectator ion.

Textbook question 7:

a. $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2\,e^-$

$Ag^+_{(aq)} + 1\,e^- \rightarrow Ag_{(s)}$

b. $Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2\,e^-$

$2\,H^+_{(aq)} + 2\,e^- \rightarrow H_{2(g)}$

Textbook question 8:

a. Bromine molecules would be formed from the bromide ions.

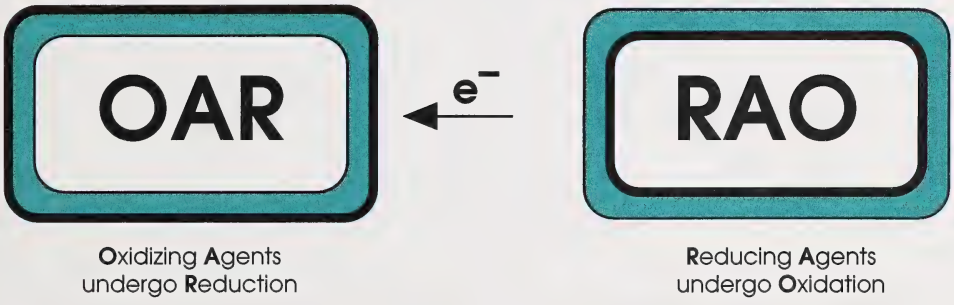
b. $2\,Br^-_{(aq)} \rightarrow Br_{2(l)} + 2\,e^-$

c. $Cl_{2(aq)} + 2\,e^- \rightarrow 2\,Cl^-_{(aq)}$

$2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,e^-$

d. The spectator ions would be sodium ions and potassium ions.

Section 1: Activity 2

1. The copper metal loses electrons and is therefore oxidized.
2. The silver ions gain electrons and therefore are reduced.
3. The copper supplies electrons so that silver ions can be reduced.
4. The silver ions gain electrons so that copper atoms can be oxidized.
5. The relationship is that the number of electrons lost must equal the number of electrons gained.
6. Copper metal is associated with LEO, since it loses electrons and is oxidized. The silver ions are associated with GER, since they gain electrons and are reduced.
7. 

8. The zinc metal is being **oxidized** by the oxidizing agent which is the **lead ion**. The zinc metal is undergoing a **loss** of electrons which is called **oxidation**. The lead ion is being **reduced** by the reducing agent which is the **zinc metal**. The lead ion is undergoing a **gain** of electrons which is called **reduction**.

9. **Textbook question 10 and question 11:**

OA – oxidizing agent **RA** – reducing agent

a. Reduction $2\left(Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)} \right)$

Oxidation $Sn_{(s)} \rightarrow Sn^{4+}_{(aq)} + 4\,e^-$

$2\,Cu^{2+}_{(aq)} + Sn_{(s)} \rightarrow 2\,Cu_{(s)} + Sn^{4+}_{(aq)}$

OA **RA**

b. Reduction $Br_{2(l)} + 2\,e^- \rightarrow 2\,Br^-_{(aq)}$

Oxidation $2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,e^-$

$Br_{2(l)} + 2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,Br^-_{(aq)}$

OA **RA**

c. Reduction $2\,H^+_{(aq)} + 2\,e^- \rightarrow H_{2(g)}$

Oxidation $Ca_{(s)} \rightarrow Ca^{2+}_{(aq)} + 2\,e^-$

$2\,H^+_{(aq)} + Ca_{(s)} \rightarrow H_{2(g)} + Ca^{2+}_{(aq)}$

OA **RA**

d. Reduction $Fe^{3+}_{(l)} + 3\,e^- \rightarrow Fe_{(l)}$

Oxidation $Al_{(s)} \rightarrow Al^{3+}_{(l)} + 3\,e^-$

$Fe^{3+}_{(l)} + Al_{(s)} \rightarrow Al^{3+}_{(l)} + Fe_{(l)}$

OA **RA**

10. Based on the assumption in the introduction, it is predicted that all the single replacement reactions will be spontaneous.

11. A brief description of the design would be as follows.

- Make sure that each of the metal strips is cleaned.
- Place a sample of each given metal in each of the given solutions, except the metal's own ion solution. It is assumed that each of the metals will not react with its own ion solution. (This can be tested, if desired.)

12. **EVIDENCE TABLE**

Metals / Metal ions	$Cu_{(s)}$	$Pb_{(s)}$	$Zn_{(s)}$	$Ag_{(s)}$ (given data)
$Cu^{2+}_{(aq)}$		S	S	NS
$Pb^{2+}_{(aq)}$	NS		S	NS
$Zn^{2+}_{(aq)}$	NS	NS		NS
$Ag^{+}_{(aq)}$	S	S	S	

13. a. The list of metal ions, from most reactive to least reactive, is as follows:

$$Ag^{+}_{(aq)},\ Cu^{2+}_{(aq)},\ Pb^{2+}_{(aq)},\ Zn^{2+}_{(aq)}$$

b. $Ag^{+}_{(aq)} + 1\,e^{-} \rightarrow Ag_{(s)}$

$Cu^{2+}_{(aq)} + 2\,e^{-} \rightarrow Cu_{(s)}$

$Pb^{2+}_{(aq)} + 2\,e^{-} \rightarrow Pb_{(s)}$

$Zn^{2+}_{(aq)} + 2\,e^{-} \rightarrow Zn_{(s)}$

14. a. The list of solid metals, from most reactive to least reactive, is as follows:

$$Zn_{(s)},\ Pb_{(s)},\ Cu_{(s)},\ Ag_{(s)}$$

b. $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^{-}$

$Pb_{(s)} \rightarrow Pb^{2+}_{(aq)} + 2\,e^{-}$

$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2\,e^{-}$

$Ag_{(s)} \rightarrow Ag^{+}_{(aq)} + 1\,e^{-}$

15. a. The metal ions are oxidizing agents.

b. The metal ions themselves were reduced.

16. a. The metals are reducing agents.

b. The metals themselves were oxidized.

17. No. Having a reducing agent and an oxidizing agent does not automatically mean a spontaneous redox reaction will take place. There has to be the right combination of oxidizing agent and reducing agent.

18. The experimental design seemed to be adequate since only qualitative tests were performed. The prediction, though, was falsified because only half of the reactions performed as predicted. This shows the need for the assumption to be either revised or thrown out.

Section 1: Activity 3

1. a. The silver ion, Ag^+, won the oxidizing agents' competition because it reacted the greatest number of times. The silver ion is called the strongest oxidizing agent.

   b. The solid zinc, $Zn_{(s)}$, won the reducing agents' competition because it reacted the greatest number of times. The zinc is named the strongest reducing agent.

2. **Textbook question 14:**

   a. The two metals that reacted spontaneously with Cu^{2+} were zinc and lead.

   b. The two metals that did not react spontaneously with Cu^{2+} were copper and silver.

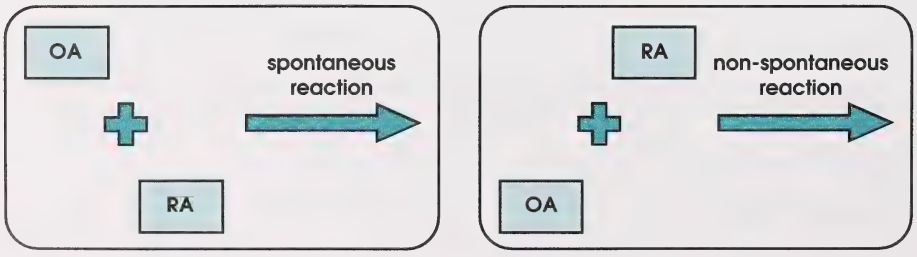
   c. The metals that reacted spontaneously with Cu^{2+} are below copper on the reduction half-reactions table.

   d. Using the table to predict the possible spontaneous reactions with the other metal ions of this investigation, the results should be as follows:

   - Ag^+ should react with $Cu_{(s)}$, $Pb_{(s)}$, and $Zn_{(s)}$.
   - Pb^{2+} should react with $Zn_{(s)}$.
   - Zn^{2+} should not react with any of the metals.

   This hypothesis was verified by the results of Investigation 12.2.

3. The redox spontaneity rule states that a spontaneous redox reaction should occur if the oxidizing agent involved is **above** the reducing agent involved on a table of reduction half-reactions.

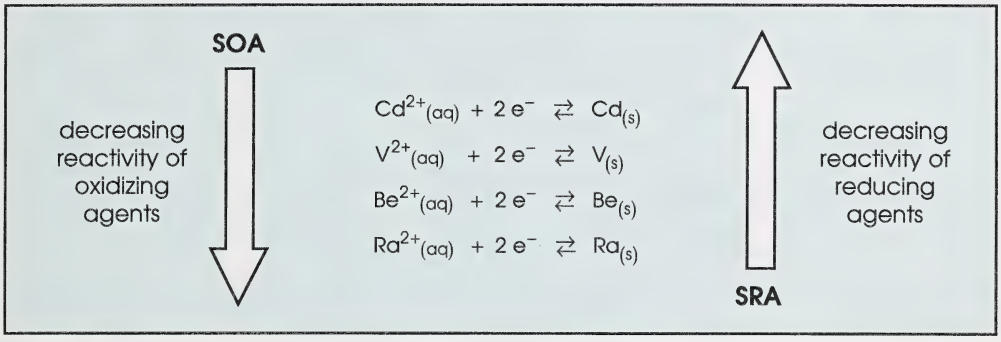
4. 

5. The various metals and metal ions, according to the given data, are as follows:

i.	$Ra^{2+}_{(aq)}$	v.	$Ra_{(s)}$
ii.	$Be^{2+}_{(aq)}$	vi.	$Be_{(s)}$
iii.	$V^{2+}_{(aq)}$	vii.	$V_{(s)}$
iv.	$Cd^{2+}_{(aq)}$	viii.	$Cd_{(s)}$

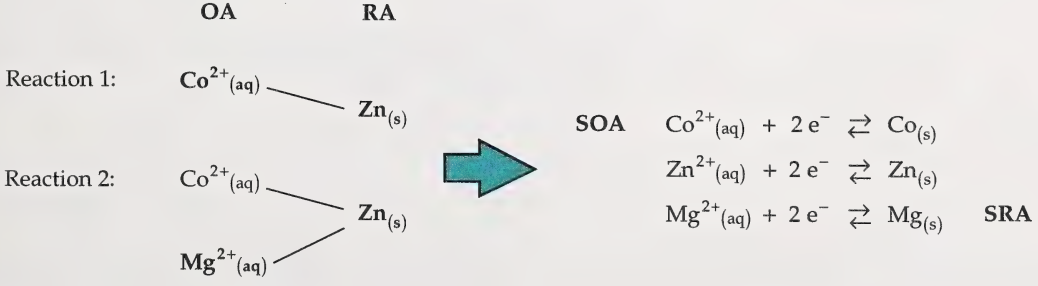
$Cd^{2+}_{(aq)}$ is the strongest oxidizing agent. $Ra_{(s)}$ is the strongest reducing agent.

6. **Redox Table**

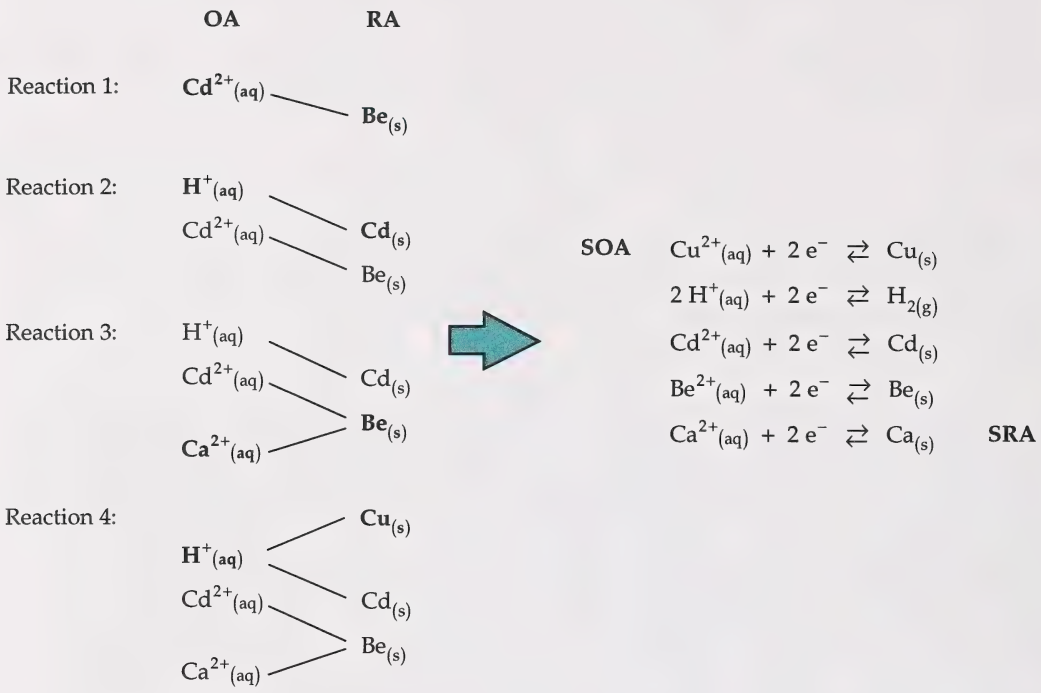

SOA
decreasing reactivity of oxidizing agents
$Cd^{2+}_{(aq)} + 2e^- \rightleftarrows Cd_{(s)}$
$V^{2+}_{(aq)} + 2e^- \rightleftarrows V_{(s)}$
$Be^{2+}_{(aq)} + 2e^- \rightleftarrows Be_{(s)}$
$Ra^{2+}_{(aq)} + 2e^- \rightleftarrows Ra_{(s)}$
SRA
decreasing reactivity of reducing agents

7. The pattern of reactions between the metals and metal ions in this lab exercise seems consistent with the evidence obtained in Investigation 12.2. Therefore, the spontaneity rule is acceptable.

8. **Textbook question 15:**



Textbook question 16:

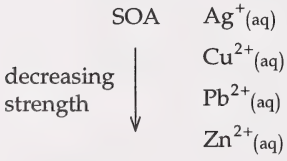


Textbook question 17:

The redox spontaneity rule is empirical because it is based on actual reaction evidence.

9. **Textbook question 19:**

The order of ions is as follows:



The order of the ions from the redox table is the same as the order developed from Investigation 12.2.

Textbook question 20:

According to the table, positive ion species (metal ions), acidic solutions, and nonmetals behave as oxidizing agents.

Textbook question 21:

According to the table, metals, nonmetal ions, and basic solutions behave as reducing agents.

Textbook question 22:

Nonmetals have nearly-filled valence electron levels which tend to make them strongly attract electrons. Since oxidizing agents also attract electrons (to be reduced), the theory of quantum mechanics is consistent with the redox table. Metals tend to lose electrons from nearly empty valence levels, which is consistent with the fact that reducing agents lose electrons (to be oxidized) as well.

Textbook question 23:

Fluorine being the most reactive nonmetal is consistent with its high position on the left-hand side of the redox table. As a strong oxidizing agent, it attracts electrons very strongly. Fluorine's observed reactivity then helps to explain its theoretical action as an oxidizing agent.

Textbook question 24:

Two metals that are found in nature as elements are gold and silver. Two metals that are never found in nature as elements are sodium and magnesium. Since gold and silver are fairly high on the reducing agents' side of the table, they are quite weak reducing agents. This means they are fairly unreactive and remain as elements. Sodium and magnesium are strong reducing agents and are quite reactive. Therefore, they react to form compounds easily, are more stable that way, and are therefore never found as elements.

Textbook question 25:

Three species that can act as either oxidizing or reducing agents are $H_2O_{(l)}$, $Sn^{2+}_{(aq)}$, and $Fe^{2+}_{(aq)}$. This is because each one can either lose electrons as reducing agents,

$$2\,H_2O_{(l)} \rightarrow O_{2(g)} + 4\,H^+_{(aq)} + 4\,e^-$$
$$Sn^{2+}_{(aq)} \rightarrow Sn^{4+}_{(aq)} + 2\,e^-$$
$$Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + 1\,e^-$$

or gain electrons as oxidizing agents,

$$2\,H_2O_{(l)} + 2\,e^- \rightarrow H_{2(g)} + 2\,OH^-_{(aq)}$$
$$Sn^{2+}_{(aq)} + 2\,e^- \rightarrow Sn_{(s)}$$
$$Fe^{2+}_{(aq)} + 2\,e^- \rightarrow Fe_{(s)}$$

Textbook question 26:

a. Nickel metal in a solution of silver ions would show a spontaneous reaction (silver ions are above nickel metal on the redox table).

b. Zinc metal in a solution of aluminum ions would be non-spontaneous (aluminum ions are below zinc metal).

c. An aqueous mixture of copper(II) ions and iodide ions would be non-spontaneous (copper(II) ions are below iodide ions).

d. Chlorine gas bubbled into a bromide ion solution would show a spontaneous reaction (chlorine is above bromide ions).

e. An aqueous mixture of copper(II) ions and tin(II) ions would show a spontaneous reaction (copper(II) ions are above tin(II) ions). (Remember, tin(II) ions are on both sides of the table.)

f. Copper metal in nitric acid would show a spontaneous reaction (the nitric acid solution, $H^+_{(aq)} + NO_3^-{}_{(aq)}$, is above copper metal).

Section 1: Activity 4

1. The strongest oxidizing agent has the greatest attraction for electrons and the strongest reducing agent has the greatest ability to donate electrons. This makes them the best competitors for the exchange of electrons.

2. **Textbook question 29:**

a. OA OA OA
$Cu^{2+}_{(aq)}$, $Pb_{(s)}$, $SO_4^{2-}{}_{(aq)}$, $H_2O_{(l)}$
RA RA RA

b. OA OA OA
$Au_{(s)}$, $H^+_{(aq)}$, $NO_3^-{}_{(aq)}$, $H_2O_{(l)}$
RA RA

c. OA OA OA OA OA OA
$K^+_{(aq)}$, $Cr_2O_7^{2-}{}_{(aq)}$, $H^+_{(aq)}$, $NO_3^-{}_{(aq)}$, $Fe^{2+}_{(aq)}$, $H_2O_{(l)}$
RA RA

(**Note:** Since you are only identifying the species themselves, and not amounts, **2** $K^+_{(aq)}$ and **2** $NO_3^-{}_{(aq)}$ is not required.)

d. OA OA OA
$Cl_{2(aq)}$, $Na^+_{(aq)}$, $OH^-_{(aq)}$, $H_2O_{(l)}$
RA RA

(Remember, Cl_2 is written as $Cl_{2(g)}$ on the right side of the redox table, but it has the same half-reaction as $Cl_{2(aq)}$.)

e. OA OA OA OA
$K^+_{(aq)}$, $MnO_4^-{}_{(aq)}$, $H^+_{(aq)}$, $Sn^{2+}_{(aq)}$, $Cl^-_{(aq)}$, $H_2O_{(l)}$
RA RA RA

f. OA OA OA
$H_2O_{(l)}$, $I_{2(s)}$, $Na^+_{(aq)}$, $SO_3^{2-}{}_{(aq)}$, $OH^-_{(aq)}$
RA RA RA RA

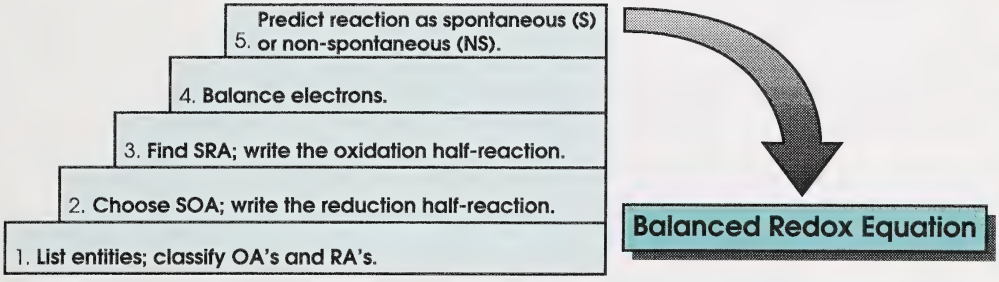
3. The strongest oxidizing agent will be the oxidizing agent that is positioned the highest on the left-hand side of the redox table. The strongest reducing agent will be the reducing agent that is positioned the lowest on the right-hand side of the redox table.

4. The redox table simply lists all the half-reactions as reduction equations because one universal way had to be chosen to write the equations. The double arrows on each equation imply that the equation can be written either way. Since both an oxidation and reduction process is needed, the equation supplying the oxidation process must be written in the reverse direction from the table.

5. The half-reaction equations must often be multiplied by a number because it is crucial that the total electrons lost and gained be balanced.

6. You use the redox spontaneity rule to decide if the equation is spontaneous. If the oxidizing agent is above the reducing agent on the redox table, then the reaction is spontaneous.

7.


5. Predict reaction as spontaneous (S) or non-spontaneous (NS).
4. Balance electrons.
3. Find SRA; write the oxidation half-reaction.
2. Choose SOA; write the reduction half-reaction.
1. List entities; classify OA's and RA's.
Balanced Redox Equation

8. **Textbook question 30:**

a. $Zn_{(s)}$, $H^+_{(aq)}$, $Cl^-_{(aq)}$, $H_2O_{(l)}$

SOA: $H^+_{(aq)}$ OA: $H_2O_{(l)}$

SRA: $Zn_{(s)}$ RA: $Cl^-_{(aq)}$ RA: $H_2O_{(l)}$

$$2\,H^+_{(aq)} + 2\,e^- \rightarrow H_{2(g)}$$
$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^-$$
$$2\,H^+_{(aq)} + Zn_{(s)} \xrightarrow{S} H_{2(g)} + Zn^{2+}_{(aq)}$$

Diagnostic Test: Hydrogen gas produces a pop when ignited.

b. $Cl_{2(g)}$, $I^-_{(aq)}$, $H_2O_{(l)}$

SOA: $Cl_{2(g)}$ OA: $H_2O_{(l)}$

SRA: $I^-_{(aq)}$ RA: $H_2O_{(l)}$

$$Cl_{2(g)} + 2\,e^- \rightarrow 2\,Cl^-_{(aq)}$$
$$2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,e^-$$
$$Cl_{2(g)} + 2\,I^-_{(aq)} \xrightarrow{S} 2\,Cl^-_{(aq)} + I_{2(s)}$$

Diagnostic Test: Iodine in solution is an orange-brown colour. If a hydrocarbon solvent is added to the solution, the iodine will turn the solvent layer purple.

c. SOA OA

$Au_{(s)}$, $H^+_{(aq)}$, $Cl^-_{(aq)}$, $H_2O_{(l)}$

RA RA RA SRA

$$2\left(2\,H^+_{(aq)} + 2\,e^- \rightarrow H_{2(g)}\right)$$

$$2\,H_2O_{(l)} \rightarrow O_{2(g)} + 4\,H^+_{(aq)} + 4\,e^-$$

$$2\,H_2O_{(l)} \xrightarrow{NS} 2\,H_{2(g)} + O_{2(g)}$$

Diagnostic Test: There is no diagnostic test since the reaction is non-spontaneous.

d. OA SOA OA

$H^+_{(aq)}$, $NO_3^-{}_{(aq)}$, $Cu_{(s)}$, $H_2O_{(l)}$

SRA RA

$$2\left(NO_3^-{}_{(aq)} + 2\,H^+_{(aq)} + e^- \rightarrow NO_{2(g)} + H_2O_{(l)}\right)$$

$$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2\,e^-$$

$$2\,NO_3^-{}_{(aq)} + 4\,H^+_{(aq)} + Cu_{(s)} \xrightarrow{S} 2\,NO_{2(g)} + 2\,H_2O_{(l)} + Cu^{2+}_{(aq)}$$

Diagnostic Test: Copper(II) ions will create a blue-coloured solution.
$NO_{2(g)}$ is a brown gas.

e. SOA OA

$Fe_{(s)}$, $O_{2(g)}$, $H_2O_{(l)}$ (**Note:** Air or wind implies O_2.)

SRA RA

$$O_{2(g)} + 2\,H_2O_{(l)} + 4\,e^- \rightarrow 4\,OH^-_{(aq)}$$

$$2\left(Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2\,e^-\right)$$

$$O_{2(g)} + 2\,H_2O_{(l)} + 2\,Fe_{(s)} \xrightarrow{S} 4\,OH^-_{(aq)} + 2\,Fe^{2+}_{(aq)}$$

Diagnostic Test: Litmus paper held in the solution near the iron will turn blue because of the hydroxide ions.

f. $Na^+_{(aq)}$, $SO_3^{2-}_{(aq)}$, $OH^-_{(aq)}$, $O_{2(g)}$, $H_2O_{(l)}$

OA: Na^+, SO_3^{2-}; SRA: SO_3^{2-}; RA: OH^-, H_2O; SOA: O_2; OA: O_2

$$O_{2(g)} + 2\,H_2O_{(l)} + 4\,e^- \rightarrow 4\,OH^-_{(aq)}$$

$$2\left(SO_3^{2-}_{(aq)} + 2\,OH^-_{(aq)} \rightarrow SO_4^{2-}_{(aq)} + H_2O_{(l)} + 2\,e^-\right)$$

$$O_{2(g)} + 2\,SO_3^{2-}_{(aq)} \xrightarrow{S} 2\,SO_4^{2-}_{(aq)}$$

Diagnostic Test: If hydrosulphuric acid, $H_2S_{(aq)}$, is added to an acidified sample of the final solution, a yellow precipitate of sulphur would likely indicate the presence of sulphate ions.

Note:

- The hydroxide ions and water molecules cancelled just like the electrons because they were present in equal amounts on both sides of the equation.
- You would not be expected to know the diagnostic test in this question for a final examination.

9. **Textbook question 31:**

SOA: H^+; OA: Sn^{2+}, H_2O; SRA: Sn^{2+}; RA: Cl^-, H_2O

$$2\,H^+_{(aq)} + 2\,e^- \rightarrow H_{2(g)}$$

$$Sn^{2+}_{(aq)} \rightarrow Sn^{4+}_{(aq)} + 2\,e^-$$

$$2\,H^+_{(aq)} + Sn^{2+}_{(aq)} \xrightarrow{NS} H_{2(g)} + Sn^{4+}_{(aq)}$$

Since the reaction is predicted as non-spontaneous, the solution should be stable.

Textbook question 32:

a. $2\,Fe_{(s)} + 3\,CuSO_{4(aq)} \rightarrow Fe_2(SO_4)_{3(aq)} + 3\,Cu_{(s)}$

$$2\,Fe_{(s)} + 3\,Cu^{2+}_{(aq)} \rightarrow 2\,Fe^{3+}_{(aq)} + 3\,Cu_{(s)}$$

← **Note:** Sulphate ions are spectator ions (do not react) and are therefore excluded from the net equation. Also, $Fe^{3+}_{(aq)}$ was chosen as a product because it is the most common iron ion.

b. SOA OA OA
$Fe_{(s)}$, $Cu^{2+}_{(aq)}$, $SO_4^{2-}_{(aq)}$, $H_2O_{(l)}$
SRA RA

$$Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$$
$$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2\,e^-$$
$$Cu^{2+}_{(aq)} + Fe_{(s)} \xrightarrow{S} Cu_{(s)} + Fe^{2+}_{(aq)}$$

c. Qualitative Design:

Check the colour of the final solution after complete reaction. If the final solution is pale green, then $Fe^{2+}_{(aq)}$ was produced. If the solution is yellow-brown, then $Fe^{3+}_{(aq)}$ was produced. (This design centres around the ion colours given on the back inside cover of *Nelson Chemistry*.)

Quantitative Design:

The masses of iron reacted and copper produced could be measured. Convert the masses to moles. If the mole to mole ratio is 1 Cu: 1 Fe, then $Fe^{2+}_{(aq)}$ was produced; if the mole to mole ratio was 3 Cu: 2 Fe, then $Fe^{3+}_{(aq)}$ was produced.

10. The prediction for Lab Exercise 12D: Testing Redox Concepts would be as follows:

OA OA SOA OA
$Sn^{2+}_{(aq)}$, $Cl^-_{(aq)}$, $H_2O_{(l)}$, $NH_4^+_{(aq)}$, $Cr_2O_7^{2-}_{(aq)}$, $H^+_{(aq)}$
SRA RA RA RA

$$Cr_2O_7^{2-}_{(aq)} + 14\,H^+_{(aq)} + 6\,e^- \rightarrow 2\,Cr^{3+}_{(aq)} + 7\,H_2O_{(l)}$$
$$3\left(Sn^{2+}_{(aq)} \rightarrow Sn^{4+}_{(aq)} + 2\,e^-\right)$$
$$Cr_2O_7^{2-}_{(aq)} + 14\,H^+_{(aq)} + 3\,Sn^{2+}_{(aq)} \xrightarrow{S} 2\,Cr^{3+}_{(aq)} + 7\,H_2O_{(l)} + 3\,Sn^{4+}_{(aq)}$$

The predicted products of the spontaneous reaction should be chromium(III) ions, water, and tin(IV) ions.

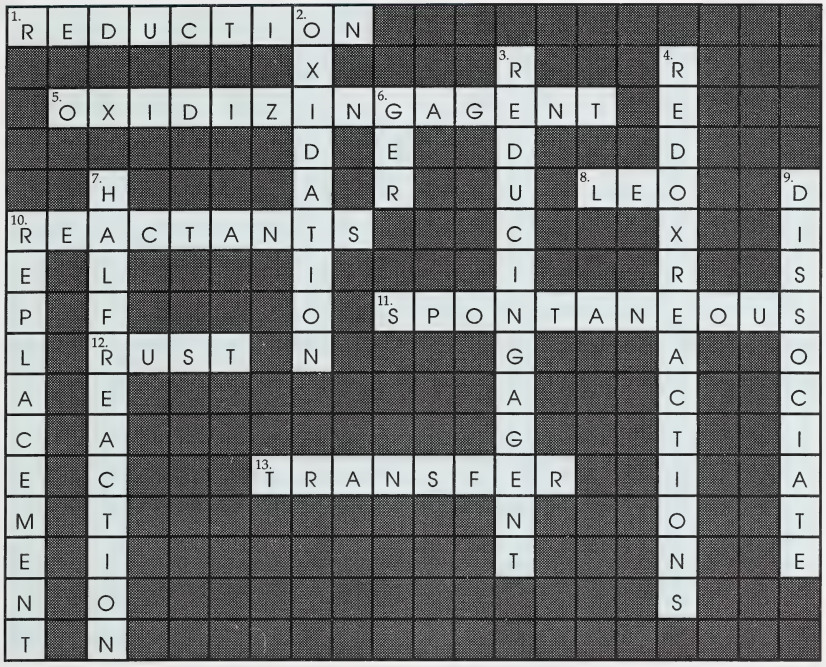
The experimental design would be as follows:

Excess amounts of tin(II) chloride are added to the ammonium dichromate solution. If the colour of the final solution is green, then chromium(III) ions have been produced.

Section 1: Follow-up Activities

Extra Help

1.



2. a. oxidation
   b. reduction
   c. reduction
   d. oxidation

3. a. $Cu^{2+}_{(aq)}$, $NO_3^-{}_{(aq)}$, $Ni_{(s)}$, $H_2O_{(l)}$

   SOA OA SRA RA

   Notice that since the solution is not acidified, $NO_3^-{}_{(aq)}$ cannot participate as an oxidizing agent.

$$Cu^{2+}_{(aq)} + 2\,e^- \rightarrow Cu_{(s)}$$
$$Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + 2\,e^-$$
$$Cu^{2+}_{(aq)} + Ni_{(s)} \xrightarrow{s} Cu_{(s)} + Ni^{2+}_{(aq)}$$

b. $\overset{\text{OA}}{K^+_{(aq)}}, \overset{\text{SOA}}{MnO_4^-{}_{(aq)}}, \overset{\text{OA}}{H^+_{(aq)}}, \overset{\text{OA}}{Cu^{2+}_{(aq)}}, \overset{\text{SRA}}{Br^-_{(aq)}}, \overset{\text{RA}}{H_2O_{(l)}}$

$$2\left(MnO_4^-{}_{(aq)} + 8\,H^+_{(aq)} + 5\,e^- \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}\right)$$
$$5\left(2\,Br^-_{(aq)} \rightarrow Br_{2(l)} + 2\,e^-\right)$$

$$2\,MnO_4^-{}_{(aq)} + 16\,H^+_{(aq)} + 10\,Br^-_{(aq)} \xrightarrow{S} 2\,Mn^{2+}_{(aq)} + 8\,H_2O_{(l)} + 5\,Br_{2(l)}$$

4. a. The list of oxidizing agents in decreasing order of strength is as follows:

$$Y^{2+}_{(aq)},\ X^{2+}_{(aq)},\ Z^{2+}_{(aq)}$$

b. The strongest agents are the ones that reacted the most. In this case the strongest oxidizing agent was $Y^{2+}_{(aq)}$ and the strongest reducing agent was $Z_{(s)}$. Therefore, the balanced equation is as follows:

$$Y^{2+}_{(aq)} + Z_{(s)} \rightarrow Y_{(s)} + Z^{2+}_{(aq)}$$

Enrichment

1. The analysis of the investigation may be as follows:

Colour Change #1:

Iodine tincture added to zinc-coated nails causes $I_{2(aq)}$ to lose its colour.

$\overset{\text{SRA}}{Zn_{(s)}}, \overset{\text{RA}}{I_{2(s)}}, \overset{\text{RA}}{H_2O_{(l)}}$ (SOA, OA)

$I_{2(s)} + 2\,e^- \rightarrow 2\,I^-_{(aq)}$ (red-brown → colourless)

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\,e^-$

(I_2 would be aqueous, but it has the same half-reaction as $I_{2(s)}$.)

Colour Change #2:

Iodine regains its colour when bleach, $Cl_{2(aq)}$, is added.

$\overset{\text{OA}}{Zn^{2+}_{(aq)}}, \overset{\text{SRA}}{I^-_{(aq)}}, \overset{\text{RA}}{Cl_{2(g)}}, \overset{\text{RA}}{H_2O_{(l)}}$ (SOA, OA)

$Cl_{2(g)} + 2\,e^- \rightarrow 2\,Cl^-_{(aq)}$ (colourless)

$2\,I^-_{(aq)} \rightarrow I_{2(s)} + 2\,e^-$ (red-brown)

(Cl_2 would be aqueous, but it has the same half-reaction as $Cl_{2(g)}$.)

The vinegar helps to clear the solution of any precipitate caused by the iodine.

2. The air you breathe could be responsible for advancing the aging process since oxygen is an oxidizing agent. As oxygen enters the body it would oxidize structures in cells such as nerve cells, therefore changing their composition. The cells would lose electrons. These changed cells would be attacked by the body's immune system causing nerve cell destruction. According to scientists at the National Research Council, oxygen can be responsible for many degenerative processes.

3. a. The reactions for the corrosion of iron exposed to the air (which contains both O_2 and water vapour) would be as follows:

SOA OA

$Fe_{(s)}$, $O_{2(g)}$, $H_2O_{(l)}$

SRA RA

$$O_{2(g)} + 2\,H_2O_{(l)} + 4\,e^- \rightarrow 4\,OH^-_{(aq)}$$

$$2\left(Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2\,e^-\right)$$

$$O_{2(g)} + 2\,H_2O_{(l)} + 2\,Fe_{(s)} \xrightarrow{S} 4\,OH^-_{(aq)} + 2\,Fe^{2+}_{(aq)}$$

The products can then combine to form one form of rust: $Fe^{2+}_{(aq)} + 2\,OH^-_{(aq)} \rightarrow 2\,Fe(OH)_{2(s)}$

$Fe^{2+}_{(aq)}$ can be further oxidized to $Fe^{3+}_{(aq)}$ and the following occurs:

$$Fe^{3+}_{(aq)} + 3\,OH^-_{(aq)} \rightarrow Fe(OH)_{3(s)}$$

red-brown rust

b. The following is only a small sample of things being done to prevent or slow down the corrosion of iron in vehicles.

- Different iron and other metal alloys which are corrosion-resistant are being used.
- Many metal parts are coated with plastic polymers or fibreglass.
- Some metal parts are galvanized (coated with zinc). The zinc itself corrodes and the oxide product provides a protective coating.

Section 2: Activity 1

1. Titration analysis involves the controlled addition of a solution of known concentration to a measured sample of unknown concentration.

2. Three possible industrial or laboratory applications of redox stoichiometry are as follows:

- A mining engineer may use redox stoichiometry to find the concentration of iron in an iron ore example.
- Chemical technicians use redox stoichiometry in determining the concentration of chemicals in products such as bleach or hydrogen peroxide.
- Lab technicians in hospitals may use redox titrations to detect trace quantities of chemicals, and redox stoichiometry to determine the chemicals' concentrations in blood and urine samples.

3. a. titrant
   b. sample
   c. endpoint
   d. colour change
   e. equivalence

4. The amounts of each reactant necessary to reach the equivalence point are related to their mole to mole ratios.

   a. The proportion of RA to OA is 2 mol $Sr_{(s)}$ to 1 mol $O_{2(g)}$; therefore, 4.0 mol $Sr_{(s)}$ to **2 mol** $O_{2(g)}$.

   b. The proportion of RA to OA is 3 mol $Mg_{(s)}$ to 1 mol $N_{2(g)}$; therefore, 6.0 mol $Mg_{(s)}$ to **2.0 mol** $N_{2(g)}$.

   c. The proportion of RA to OA is 1 mol $Zn_{(s)}$ to 1 mol $I_{2(s)}$; therefore, 2.5 mol $Zn_{(s)}$ to **2.5 mol** $I_{2(s)}$.

   d. The proportion of RA to OA is 4 mol $Fe_{(s)}$ to 3 mol $O_{2(g)}$; therefore, 12.0 mol $Fe_{(s)}$ to **9.0 mol** $O_{2(g)}$.

   e. The proportion of RA to OA is 1 mol $Si_{(s)}$ to 2 mol $F_{2(g)}$; therefore, 3.5 mol $Si_{(s)}$ to **7.0 mol** $F_{2(g)}$.

5. a. $MnO_4^-{}_{(aq)}$ **(purple)** $\rightarrow$ $Mn^{2+}{}_{(aq)}$ **(pale pink)**

   b. $Cr_2O_7^{2-}{}_{(aq)}$ **(orange)** $\rightarrow$ $Cr^{3+}{}_{(aq)}$ **(green)**

   c. $Fe^{3+}{}_{(aq)}$ **(yellow-brown)** $\rightarrow$ $Fe^{2+}{}_{(aq)}$ **(pale green)**

6. An endpoint in a titration is the point reached when there is an observable abrupt change (like colour) in the solution property. An equivalence point in a titration is the theoretical point reached when chemically equivalent amounts of all reactants are present in the solution. The equivalence point is usually reflected by the volume of titrant added at the endpoint.

7. Potassium permanganate is not used as a primary standard because it reacts with impurities in the water and with the water itself. It does not maintain a stable concentration.

8. Acidified tin(II) chloride solution is used as the primary standard.

9. The oxidizing agent is $MnO_4^-{}_{(aq)} + H^+{}_{(aq)}$; the reducing agent is $Sn^{2+}{}_{(aq)}$. Checking the Redox Half-Reactions table on page 552 of *Nelson Chemistry*, the oxidizing agent is above the reducing agent, thereby confirming a spontaneous reaction.

10. The endpoint of the first trial was overshot (as evidenced by the dark pink colour) and is therefore not an accurate result. This trial result was left out of the average calculation.

11. The appearance of a light pink colour in the solution indicated the endpoint.

12. $C_{(MnO_4^-)} = 10.00 \cancel{mL\ Sn^{2+}} \times \dfrac{0.0500\ \cancel{mol\ Sn^{2+}}}{\cancel{L\ Sn^{2+}}} \times \dfrac{2\ mol\ \cancel{MnO_4^-}}{5\ \cancel{mol\ Sn^{2+}}} \times \dfrac{1}{16.8\ \cancel{mL}\ \cancel{MnO_4^-}}$

$= 0.0119\ mol/L$

13. a. $\%\ difference = \dfrac{|0.0119\ mol/L - 0.0125\ mol/L|}{|0.0125\ mol/L|} \times 100\%$

$= 4.8\%$ The percent difference is 4.8%.

b. Yes. The relatively high percent difference indicates that the $MnO_4^-(aq)$ solution cannot be used without standardization.

14. a. The analysis of the tin investigation report is as follows:

$$\text{average volume of } KMnO_{4(aq)} = \frac{(12.4\ mL + 12.3\ mL + 12.5\ mL)}{3}$$

$$= 12.4\ mL$$

OA OA SOA OA OA

$Sn^{2+}(aq)$, $F^-(aq)$, $K^+(aq)$, $MnO_4^-(aq)$, $H^+(aq)$, $H_2O_{(l)}$

SRA RA RA

$$2\left(MnO_4^-(aq) + 8\,H^+(aq) + 5\,e^- \rightarrow Mn^{2+}(aq) + 4\,H_2O_{(l)}\right)$$

$$5\left(Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2\,e^-\right)$$

$$2\,MnO_4^-(aq) + 16\,H^+(aq) + 5\,Sn^{2+}(aq) \xrightarrow{S} 2\,Mn^{2+}(aq) + 8\,H_2O_{(l)} + 5\,Sn^{4+}(aq)$$

$0.0832\ mol/L$ $10.00\ mL$

$12.4\ mL$ $C = ?$

$$C_{(Sn^{2+})} = 12.4\ \cancel{mL\ MnO_4^-} \times \frac{0.0832\ \cancel{mol\ MnO_4^-}}{\cancel{L\ MnO_4^-}} \times \frac{5\ mol\ \cancel{Sn^{2+}}}{2\ \cancel{mol\ MnO_4^-}} \times \frac{1}{10.00\ \cancel{mL}\ \cancel{Sn^{2+}}}$$

$= 0.258\ mol/L$

The concentration of $Sn^{2+}(aq)$ is 0.258 mol/L.

b. The analysis of the chromium investigation report is as follows:

$$\text{average volume of } K_2Cr_2O_{7(aq)} = 17.4\ mL$$

OA OA SOA OA OA

$Cr^{2+}(aq)$, $K^+(aq)$, $Cr_2O_7^{2-}(aq)$, $H^+(aq)$, $H_2O_{(l)}$

SRA RA

$$Cr_2O_7^{2-}(aq) + 14\,H^+(aq) + 6\,e^- \rightarrow 2\,Cr^{3+}(aq) + 7\,H_2O_{(l)}$$
$$6\left(Cr^{2+}(aq) \rightarrow Cr^{3+}(aq) + 1\,e^-\right)$$

$$Cr_2O_7^{2-}(aq) + 14\,H^+(aq) + 6\,Cr^{2+}(aq) \xrightarrow{S} 8\,Cr^{3+}(aq) + 7\,H_2O_{(l)}$$

0.125 mol/L	10.00 mL
17.4 mL	C = ?

$$C_{(Cr^{2+})} = 17.4\ \cancel{mL\ Cr_2O_7^{2-}} \times \frac{0.125\ \cancel{mol\ Cr_2O_7^{2-}}}{\cancel{L\ Cr_2O_7^{2-}}} \times \frac{6\ mol\ \cancel{Cr^{2+}}}{1\ \cancel{mol\ Cr_2O_7^{2-}}} \times \frac{1}{10.00\ \cancel{mL}\ \cancel{Cr^{2+}}}$$
$$= 1.31\ mol/L$$

The concentration of $Cr^{2+}(aq)$ is 1.31 mol/L.

15. The analysis and evaluation of the iron investigation report is as follows:

$$\text{average volume of } Ce^{4+}(aq) = \frac{(15.1 + 15.0 + 14.9 + 14.9)\ mL}{4}$$
$$= 15.0\ mL\ (14.975\ mL)$$

OA **SOA** **OA**

$Fe^{2+}(aq)$, $Ce^{4+}(aq)$, $H_2O_{(l)}$

SRA **RA**

$$Ce^{4+}(aq) + 1\,e^- \rightarrow Ce^{3+}(aq)$$
$$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + 1\,e^-$$

$$Ce^{4+}(aq) + Fe^{2+}(aq) \xrightarrow{S^*} Ce^{3+}(aq) + Fe^{3+}(aq)$$

0.125 mol/L	25.0 mL
14.975 mL	C = ?

(***Note:** The reaction is assumed spontaneous since there was an endpoint reached.)

$$C_{(Fe^{2+})} = 14.975\ \cancel{mL\ Ce^{4+}} \times \frac{0.125\ \cancel{mol\ Ce^{4+}}}{\cancel{L\ Ce^{4+}}} \times \frac{1\ mol\ \cancel{Fe^{2+}}}{1\ \cancel{mol\ Ce^{4+}}} \times \frac{1}{25.0\ \cancel{mL}\ \cancel{Fe^{2+}}}$$
$$= 0.074\ 875\ mol/L$$
$$= 74.9\ mmol/L$$

The concentration of $Fe^{2+}(aq)$ was 74.9 mmol/L. Since the predicted value for the concentration of $Fe^{2+}(aq)$ was 80.0 mmol/L, the percent difference is as follows:

$$\%\ \text{difference} = \frac{|74.9\ mmol/L - 80.0\ mmol/L|}{|80.0\ mmol/L|} \times 100\%$$
$$= 6.4\%$$

For precise processes such as titration analysis, many sources quote a percent difference or error of 5% as a guideline for valid results. Since the percent difference is greater than 5% in this investigation, the prediction cannot be confirmed with confidence. The percent difference may be accounted for by uncertainties in the experiment, but this is in doubt. The process seems acceptable because of the consistency of results; also, the metallurgical process probably has some built-in tolerances for error.

16. The percent concentration on consumer brands of hydrogen peroxide is usually listed as 3% (3 g/100 mL).

17. The mass of $FeSO_4 \bullet (NH_4)_2SO_4 \bullet 6H_2O_{(s)}$ required is 1.96 g.

$$m = 0.1000\ \text{L} \times \frac{0.0500\ \text{mol}}{\text{L}} \times \frac{392.19\ \text{g}}{\text{mol}}$$
$$= 1.96\ \text{g}$$

18. The following chart illustrates sample data:

Standardization of $KMnO_{4(aq)}$ Solution (volume of $KMnO_{4(aq)}$ required to react with 10.00 mL of 0.0500 mol/L $Fe^{2+}{}_{(aq)}$)				
Trial	**1**	**2**	**3**	**4**
Final buret reading (mL)	29.2	42.7	13.1	26.2
Initial buret reading (mL)	16.2	29.2	0.2	13.1
Endpoint colour	pink	red	pink	pink

The value for Trial 2 will not be used in the average because the endpoint was overrun.

19. The following chart illustrates sample data:

Titration of $H_2O_{2(aq)}$ Solution (volume of standardized $KMnO_{4(aq)}$ required to titrate 10.00 mL of $H_2O_{2(aq)}$)			
Trial	**1**	**2**	**3**
Final buret reading (mL)	12.3	23.5	11.4
Initial buret reading (mL)	1.2	12.3	0.2
Endpoint colour	pink	pink	pink

20. Refer to the answer for Section 2: Activity 1, question 18. The chart illustrates the given data.

21. Refer to the answer for Section 2: Activity 1, question 19. The chart illustrates the given data.

22. The $H_2SO_{4(aq)}$ is required in the overall reaction to supply $H^+{}_{(aq)}$ ions to $MnO_4{}^-{}_{(aq)}$ to serve as the strong oxidizing agent. The $H_2SO_{4(aq)}$ is not mixed directly with the $KMnO_{4(aq)}$ to begin with because of the unstable nature of the $KMnO_{4(aq)}$.

23. Stage 1 – Standardization

$$\text{average volume of } KMnO_{4(aq)} = \frac{(13.0 + 12.9 + 13.1)\ mL}{3}$$
$$= 13.0\ mL$$

24. The concentration of the $KMnO_{4(aq)}$ is 0.007 69 mol/L.

OA OA OA OA SOA OA

$Fe^{2+}_{(aq)}$, $NH_4^+{}_{(aq)}$, $SO_4^{2-}{}_{(aq)}$, $H_2O_{(l)}$, $K^+_{(aq)}$, $MnO_4^-{}_{(aq)}$, $H^+_{(aq)}$, $HSO_4^-{}_{(aq)}$ ←

SRA RA

Note: $H_2SO_{4(aq)}$ is only strong for ionizing the first hydrogen – therefore it is written as shown.

$$MnO_4^-{}_{(aq)} + 8\,H^+_{(aq)} + 5\,e^- \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}$$
$$5\left(Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + 1\,e^-\right)$$

$$\xrightarrow{s} MnO_4^-{}_{(aq)} + 8\,H^+_{(aq)} + 5\,Fe^{2+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)} + 5\,Fe^{3+}_{(aq)}$$

13.0 mL	0.0500 mol/L
C = ?	10.00 mL

$$C_{(MnO_4^-)} = 10.00\ \cancel{mL\ Fe^{2+}} \times \frac{0.0500\ \cancel{mol\ Fe^{2+}}}{\cancel{L\ Fe^{2+}}} \times \frac{1\ mol\ \cancel{MnO_4^-}}{5\ \cancel{mol\ Fe^{2+}}} \times \frac{1}{13.0\ \cancel{mL\ MnO_4^-}}$$
$$= 0.007\ 69\ mol/L\ (0.007\ 6923\ mol/L)$$

25. Stage 2 – Titration

$$\text{average volume of } KMnO_{4(aq)} = \frac{(11.1 + 11.2 + 11.2)\ mL}{3}$$
$$= 11.2\ mL\ (11.167\ mL)$$

26. The concentration of the $H_2O_{2(aq)}$ is 0.0215 mol/L.

OA SOA OA OA

$H_2O_{2(aq)}$, $HSO_4^-{}_{(aq)}$, $H^+_{(aq)}$, $MnO_4^-{}_{(aq)}$, $K^+_{(aq)}$, $H_2O_{(l)}$

SRA RA

Note: Even though the redox table lists the state of H_2O_2 as liquid $_{(l)}$, the half-reaction of the aqueous $_{(aq)}$ H_2O_2 will be the same.

$$2\left(MnO_4^-{}_{(aq)} + 8\,H^+_{(aq)} + 5\,e^- \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}\right)$$
$$5\left(H_2O_{2(aq)} \rightarrow O_{2(g)} + 2\,H^+_{(aq)} + 2\,e^-\right) \leftarrow$$

$$\xrightarrow{s} 2\,MnO_4^-{}_{(aq)} + 6\,H^+_{(aq)} + 5\,H_2O_{2(aq)} \rightarrow 2\,Mn^{2+}_{(aq)} + 8\,H_2O_{(l)} + 5\,O_{2(g)}$$

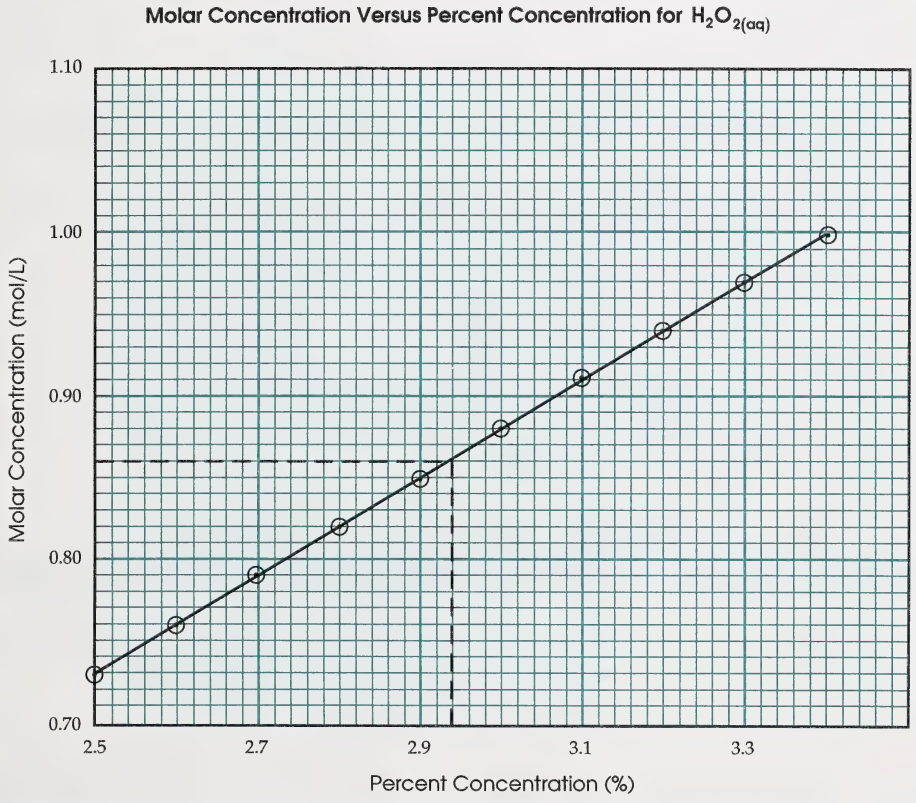
0.007 6923 mol/L	10.00 mL
11.167 mL	C = ?

$$C_{(H_2O_2)} = 11.167\ \cancel{mL\ MnO_4} \times \frac{0.007\ 6923\ \cancel{mol\ MnO_4}}{\cancel{L\ MnO_4}} \times \frac{5\ mol\ \cancel{H_2O_2}}{2\ \cancel{mol\ MnO_4}} \times \frac{1}{10.00\ \cancel{mL}\ \cancel{H_2O_2}}$$
$$= 0.0215\ mol/L$$

27. The molar concentration of H_2O_2 is 0.86 mol/L.

$$C_{(H_2O_2)} = 0.0215\ mol/L \times 40$$
$$= 0.86\ mol/L$$

28. a. The graph of percent concentration and molar concentration of $H_2O_{2(aq)}$ is as follows.



b. The interpolated value for the percent concentration is 2.94%, as marked on the graph.

29. The evaluation of the investigation is as follows:

a. $$\text{percent difference} = \frac{|2.94\% - 3\%|}{|3\%|} \times 100\%$$
$$= 2\%$$

b. Based on the low percent difference, the prediction is judged to be verified. The small variation in percentages could be attributed to limits in equipment precision or uncertainty in measurements.

c. The titration design is judged acceptable considering the ability to collect adequate amounts of data and the consistent results obtained.

Section 2: Activity 2

1. Numbers are much easier to keep track of in the computer systems and other areas of institutions.

2. The method of electron bookkeeping was developed to describe oxidation and reduction in molecules and polyatomic (complex) ions. The method allows chemists to keep track of the loss and gain of electrons.

3. a. The oxidation state of an atom is the apparent net electric charge that an atom would have if the electrons in a covalent bond were completely controlled by the more electronegative atom.

b. No, the oxidation state of an atom represents an imaginary charge, not a real charge.

4. Ion charges and oxidation numbers are distinguished as follows:

Ion Charges

$\boxed{2^-}$ ← number before charge

Oxidation Numbers

$\boxed{-2}$ ← charge before number

5. **ASSIGNING COMMON OXIDATION NUMBERS**

Atom or Ion	Oxidation Number	Examples
hydrogen in compounds	+1	**H in H_2S is +1; H in NaOH is +1**
oxygen in compounds	–2	**O in CO_2 is –2; O in SO_3^{2-} is –2**
all atoms in elements	**0**	Mg is 0; N in N_2 is 0
all ions	same size as charge on the ion	**Ba^{2+} is +2; P^{3-} is –3; PO_4^{3-} is –3**

6. The exceptions for oxygen and hydrogen are as follows:

- oxygen – The oxidation number for O in peroxides is –1, for example $\overset{+1}{H_2}\overset{\boxed{-1}}{O_2}$.
  – The oxidation number for O in OF_2 is +2.

- hydrogen – The oxidation number for H in hydrides is –1 when the partner atom is less electronegative, for example, $\overset{+1}{Na}\overset{\boxed{-1}}{H}$.

7. The oxidation number for *N* in NH_3 is assigned as follows:

- Each H in NH_3 is +1. → $\overset{+1}{N}H_3$
- Since there are three H's, the total positive count for NH_3 is +3. → $\overset{+1}{N}H_3$ $3(+1)$
- The total charge on NH_3 is zero, so the total negative count on NH_3 must be –3. → $\overset{+1}{N}H_3$. $-3+3(+1)=0$
- There is only one N atom to supply the –3, so the oxidation number of N must be –3. → $\overset{\boxed{-3}\ +1}{N}H_3$ $1(-3)+3(+1)=0$

8. The oxidation number for *S* in $S_2O_3^{2-}$ is assigned as follows:

- The oxidation number for O is –2. → $S_2\overset{-2}{O}_3^{2-}$
- Since there are three O's, the total negative count for $S_2O_3^{2-}$ is -6. → $S_2\overset{-2}{O}_3^{2-}$ $3(-2)$
- The oxidation number on $S_2O_3^{2-}$ is –2, so the total positive count must be $+4$. → $S_2\overset{-2}{O}_3^{2-}$ $+4+3(-2)=-2$
- With two S atoms to supply the $+4$, the oxidation number of each S must be +2. → $\overset{\boxed{+2}\ -2}{S_2}O_3^{2-}$ $2(+2)+3(-2)=-2$

9. **Textbook question 39:**

a. $S = +4$ $\overset{+4\ -2}{S}O_2$

$1(x)+2(-2)=0$

$x=+4$

b. $Cl = +7$ $\overset{+1\ +7\ -2}{H}\overset{+7}{Cl}O_4$

$1(+1)+1(x)+4(-2)=0$

$x=+7$

c. $S = 6$ $\overset{+6}{S}\overset{-2}{O_4}^{2-}$

$1(x) + 4(-2) = -2$

$x = +6$

d. $Cr = +6$ $\overset{+6}{Cr_2}\overset{-2}{O_7}^{2-}$

$2(x) + 7(-2) = -2$

$x = +6$

e. $I = -1$ $\overset{+2}{Mg}\overset{-1}{I_2}$

$1(+2) + 2(x) = 0$

$x = -1$

Textbook question 40:

a. $N = +1$ $\overset{+1}{N_2}\overset{-2}{O}$

$2(x) + 1(-2) = 0$

$x = +1$

b. $N = +2$ $\overset{+2}{N}\overset{-2}{O}$

$1(x) + 1(-2) = 0$

$x = +2$

c. $N = +4$ $\overset{+4}{N}\overset{-2}{O_2}$

$1(x) + 2(-2) = 0$

$x = +4$

d. $N = -3$ $\overset{-3}{N}\overset{+1}{H_3}$

$1(x) + 3(+1) = 0$

$x = -3$

e. $N = -2$ $\overset{-2}{N_2}\overset{+1}{H_4}$

$2(x) + 4(+1) = 0$

$x = -2$

f. $N = +5$ $\overset{+1}{Na}\overset{+5}{N}\overset{-2}{O_3}$

$1(+1) + 1(x) + 3(-2) = 0$

$x = +5$

g. $N = 0$ $\overset{0}{N_2}$

$2(x) = 0$

$x = 0$

h. $N = -3$ $\overset{-3}{N}\overset{+1}{H_4}\overset{-1}{Cl}$

$1(x) + 4(+1) + 1(-1) = 0$

$x = -3$

Textbook question 41:

a. $C = 0$ $\overset{0}{C}$

b. $C = 0$ $\overset{0}{C_6}\overset{+1}{H_{12}}\overset{-2}{O_6}$

$6(x) + 12(+1) + 6(-2) = 0$

$x = 0$

c. $C = +4$ $\overset{+1}{Na_2}\overset{+4}{C}\overset{-2}{O_3}$

$2(+1) + 1(x) + 3(-2) = 0$

$x = +4$

d. $C = +2$ $\overset{+2}{C}\overset{-2}{O}$

$1(x) + 1(-2) = 0$

$x = +2$

10. a. $\overset{+2}{Fe}\,\overset{+4}{S}\,\overset{-2}{O_3}$

$\overset{+2}{Fe}^{2+}$ $\qquad$ $\overset{+4}{S}\,\overset{-2}{O_3}^{2-}$

$1(x) = +2$ $\qquad$ $1(x) + 3(-2) = -2$

$x = +2$ $\qquad$ $x = +4$

b. $\overset{+1}{Ag}\,\overset{+5}{N}\,\overset{-2}{O_3}$

$\overset{+1}{Ag}^{+}$ $\qquad$ $\overset{+5}{N}\,\overset{-2}{O_3}^{-}$

$1(x) = +1$ $\qquad$ $1(x) + 3(-2) = -1$

$x = 1$ $\qquad$ $x = +5$

c. $\overset{+2}{Pb_3}\left(\overset{+5}{P}\,\overset{-2}{O_4}\right)_2$

$\overset{+2}{Pb}^{2+}$ $\qquad$ $\overset{+5}{P}\,\overset{-2}{O_4}^{3-}$

$1(x) = +2$ $\qquad$ $1(x) + 4(-2) = -3$

$x = +2$ $\qquad$ $x = +5$

d. $\overset{+3}{Al}\left(\overset{+1}{H_2}\,\overset{+5}{P}\,\overset{-2}{O_4}\right)_3$

$\overset{+3}{Al}^{3+}$ $\qquad$ $\overset{+1}{H_2}\,\overset{+5}{P}\,\overset{-2}{O_4}^{-}$

$1(x) = +3$ $\qquad$ $2(+1) + 1(x) + 4(-2) = -1$

$x = +3$ $\qquad$ $x = +5$

e. $\overset{+1}{K_2}\,\overset{+6}{Cr_2}\,\overset{-2}{O_7}$

$\overset{+1}{K}^{+}$ $\qquad$ $\overset{+6}{Cr_2}\,\overset{-2}{O_7}^{2-}$

$1(x) = +1$ $\qquad$ $2(x) + 7(-2) = -2$

$x = +1$ $\qquad$ $x = +6$

11. **Textbook question 43:**

a. $\overset{0}{Cu_{(s)}} + 2\ \overset{+1}{Ag}\,\overset{+5}{N}\,\overset{-2}{O_{3(aq)}} \rightarrow 2\ \overset{0}{Ag_{(s)}} + \overset{+2}{Cu}\left(\overset{+5}{N}\,\overset{-2}{O_3}\right)_{2(aq)}$

reduction (gain of $1\ e^-$)

RA OA

oxidation (loss of $2\ e^-$)

The equation represents a redox reaction.

Notes

- The coefficients in front of the chemical substances have no bearing on the oxidation numbers, but the subscript numbers in the formulas do affect oxidation numbers.

Example: $\boxed{2}\ AgNO_{\boxed{3}(aq)}$

does not affect oxidation numbers

does affect oxidation numbers

- The descriptions given on the arrows (gain of $1\ e^-$, loss of $2\ e^-$) are **per atom** of the element involved.

b. $\overset{+2}{Pb}(\overset{+5}{N}\overset{-2}{O_3})_{2(aq)} + 2\,\overset{+1}{K}\overset{-1}{I}_{(aq)} \rightarrow \overset{+2}{Pb}\overset{-1}{I}_{2(s)} + 2\,\overset{+1}{K}\overset{+5}{N}\overset{-2}{O}_{3(aq)}$

Oxidation numbers do not change; the equation does not represent a redox reaction.

c. reduction (gain of $1\,e^-$)

$\overset{0}{Cl}_{2(aq)} + 2\,\overset{+1}{K}\overset{-1}{I}_{(aq)} \rightarrow \overset{0}{I}_{2(s)} + 2\,\overset{+1}{K}\overset{-1}{Cl}_{(aq)}$

OA RA

oxidation (loss of $1\,e^-$)

This equation represents a redox reaction.

d. reduction (gain of $1\,e^-$)

$2\,\overset{+1}{Na}\overset{-1}{Cl}_{(l)} \rightarrow 2\,\overset{0}{Na}_{(l)} + \overset{0}{Cl}_{2(g)}$

OA RA

oxidation (loss of $1\,e^-$)

This equation represents a redox reaction.

e. $\overset{+1}{H}\overset{-1}{Cl}_{(aq)} + \overset{+1}{Na}\overset{-2}{O}\overset{+1}{H}_{(aq)} \rightarrow \overset{+1}{H}\overset{-2}{O}\overset{+1}{H}_{(l)} + \overset{+1}{Na}\overset{-1}{Cl}_{(aq)}$

Oxidation numbers do not change; this equation does not represent a redox reaction.

f. reduction (gain of $1\,e^-$)

$2\,\overset{0}{Al}_{(s)} + 3\,\overset{0}{Cl}_{2(g)} \rightarrow 2\,\overset{+3}{Al}\overset{-1}{Cl}_{3(s)}$

RA OA

oxidation (loss of $3\,e^-$)

This equation represents a redox reaction.

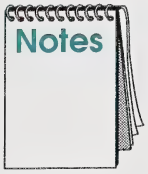
g. reduction (gain of $2\,e^-$)

$2\,\overset{-\frac{5}{2}}{C}_4\overset{+1}{H}_{10(g)} + 13\,\overset{0}{O}_{2(g)} \rightarrow 8\,\overset{+4}{C}\overset{-2}{O}_{2(g)} + 10\,\overset{+1}{H}_2\overset{-2}{O}_{(l)}$

RA OA

oxidation (loss of $6.5\,e^-$)

This equation represents a redox reaction.



As you see in this question, oxidation numbers can be fractional $\left(-\frac{5}{2}\right)$. This further points out the difference between oxidation numbers and ion charges – oxidation numbers represent the *apparent* oxidation state of an atom, not a real value.

h. reduction (gain of 5 e^-)

$$5\ \overset{-2\ +1\ \ -2\ +1}{CH_3OH_{(l)}} + 2\ \overset{+7\ -2}{MnO_4^-{}_{(aq)}} + 6\ \overset{+1}{H^+_{(aq)}} \rightarrow 5\ \overset{0\ +1\ -2}{CH_2O_{(l)}} + 2\ \overset{+2}{Mn^{2+}_{(aq)}} + 8\ \overset{+1\ -2}{H_2O_{(l)}}$$

RA OA

oxidation (loss of 2 e^-)

This equation represents a redox reaction.

12. a. reduction (gain of 4 e^-)

$$\text{energy} + 6\ \overset{+4\ -2}{CO_{2(g)}} + 6\ \overset{+1\ -2}{H_2O_{(l)}} \rightarrow \overset{0\ +1\ -2}{C_6H_{12}O_{6(s)}} + 6\ \overset{0}{O_{2(g)}}$$

OA RA RA

oxidation (loss of 2 e^-)

b. Since the oxidizing agent is carbon, $CO_{2(g)}$ and $C_6H_{12}O_{6(s)}$ would be in the reduction half-reaction equation.

$$CO_{2(g)} + ? \rightarrow C_6H_{12}O_{6(s)} + ?$$

Note: The next activity will explain how this half-reaction equation would be completed.

c. Since the reducing agent is oxygen, $CO_{2(g)}$, $H_2O_{(l)}$, and $O_{2(g)}$ would be in the oxidation half-reaction equation.

$$? + CO_{2(g)} + H_2O_{(l)} \rightarrow O_{2(g)} + ?$$

In summary,

- The oxidation half-reaction will contain any reactant(s) containing the reducing agent and any product(s) containing the same element whose oxidation number has increased.
- The reduction half-reaction will contain any reactant(s) containing the oxidizing agent and any product(s) containing the same element whose oxidation number has decreased.

d. $6\,CO_{2(g)} + 6\,H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + 6\,O_{2(g)} \quad \Delta H°_r = ?$

$6(-393.5) \quad 6(-285.8) \quad 1(-1273.1) \quad 6(0)$

$$\Delta H°_r = \sum nH°_{fp} - \sum nH°_{fr}$$

$$= \left(1\text{ mol} \times \frac{-1273.1\text{ kJ}}{\text{mol}} + 6\text{ mol} \times \frac{0\text{ kJ}}{\text{mol}}\right) - \left(6\text{ mol} \times \frac{-393.5\text{ kJ}}{\text{mol}} + 6\text{ mol} \times \frac{-285.8\text{ kJ}}{\text{mol}}\right)$$

$$= +2802.7\text{ kJ}$$

The enthalpy value for the photosynthesis reaction would be +2802.7 kJ. The photosynthesis reaction is endothermic. The energy required to carry out this reaction comes from the sun.

13. a. reduction (gain of $2\,e^-$)

$\overset{-4\,+1}{C}H_{4(g)} + 2\,\overset{0}{O}_{2(g)} \rightarrow \overset{+4\,-2}{C}O_{2(g)} + 2\,\overset{+1}{H}_2\overset{-2}{O}_{(g)}$ + energy

RA OA

oxidation (loss of $8\,e^-$)

b. Since the oxidizing agent is oxygen, $O_{2(g)}$, $CO_{2(g)}$, and $H_2O_{(g)}$ would be in the reduction half-reaction.

$$? + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)} + ?$$

c. Since carbon is the reducing agent, $CH_{4(g)}$ and $CO_{2(g)}$ would be in the oxidation half-reaction.

$$? + CH_{4(g)} \rightarrow CO_{2(g)} + ?$$

d. $CH_{4(g)} + 2\,O_{2(g)} \rightarrow CO_{2(g)} + 2\,H_2O_{(g)} \quad \Delta H°_r = ?$

$1(-74.4) \quad 2(0) \quad 1(-393.5) \quad 2(-241.8)$

$$\Delta H°_r = \sum nH°_{fp} - \sum nH°_{fr}$$

$$= \left(1\text{ mol} \times \frac{-393.5\text{ kJ}}{\text{mol}} + 2\text{ mol} \times \frac{-241.8\text{ kJ}}{\text{mol}}\right) - \left(1\text{ mol} \times \frac{-74.4\text{ kJ}}{\text{mol}} + 2\text{ mol} \times \frac{0\text{ kJ}}{\text{mol}}\right)$$

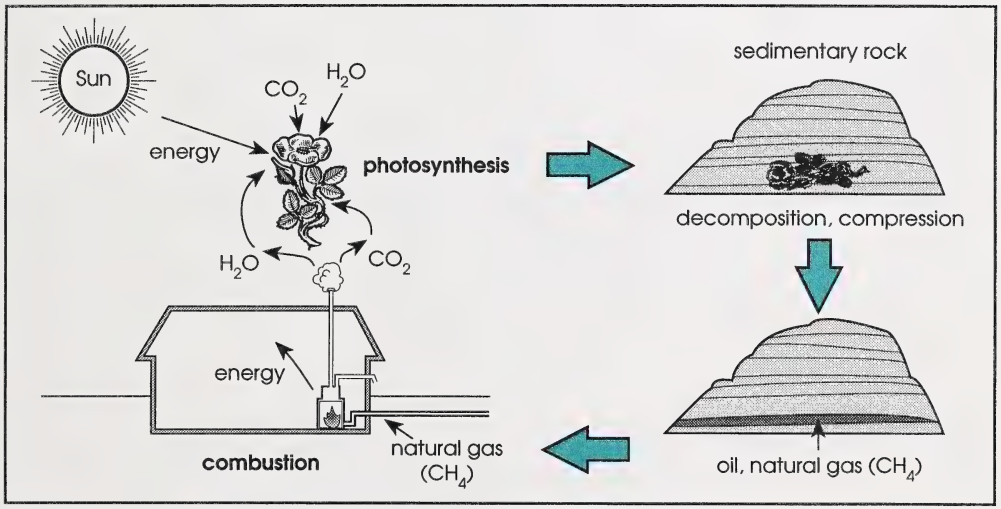
$$= -802.7\text{ kJ}$$

The enthalpy change for the combustion of methane is –802.7 kJ.

14.

PHOTOSYNTHESIS AND METHANE COMBUSTION	
Similarities	**Differences**
• Both reactions involve carbon dioxide, oxygen, and water. • Both involve organic molecules. • Both have their origins in plants. • Both are redox reactions.	• The photosynthesis reaction is endothermic whereas the methane combustion is exothermic. • Photosynthesis reacts $CO_{2(g)}$ and $H_2O_{(l)}$ whereas methane combustion produces $CO_{2(g)}$ and $H_2O_{(g)}$.

The two reactions are connected in the following manner:



Section 2: Activity 3

1. a. **SOA** **OA**
$Br_{2(l)}$, $H_2Te_{(aq)}$, $H_2O_{(l)}$
SRA **RA**

$$Br_{2(l)} + 2\,e^- \rightarrow 2\,Br^-_{(aq)}$$
$$H_2Te_{(aq)} \rightarrow Te_{(s)} + 2\,H^+_{(aq)} + 2\,e^-$$
$$Br_{2(l)} + H_2Te_{(aq)} \xrightarrow{S} 2\,Br^-_{(aq)} + Te_{(s)} + 2\,H^+_{(aq)}$$

b. Other evidence to support a spontaneous reaction prediction would be a precipitate or solid forming $\left(Te_{(s)}\right)$.

2. The redox table does not allow you to balance all redox equations because not all redox reactions will have the appropriate half-reaction equations on the redox table to predict and balance their overall equations. Also, more complicated redox equations are too difficult to balance by half-reactions.

3. The important assumption in both methods of balancing redox equations is that the **total electrons gained equals the total electrons lost.**

4. **Textbook question 46:**

a. $\overset{0}{H_{2(g)}} + \overset{+3}{Fe_2}\overset{-2}{O_{3(aq)}} \rightarrow 2\ \overset{+2}{Fe}\overset{-2}{O_{(s)}} + \overset{+1}{H_2}\overset{-2}{O_{(g)}}$

$1\ e^- / H$ $1\ e^- / Fe$

$2\ e^- / H_2$ $2\ e^- / Fe_2O_3$

(×1) (×1)

b. $2\ \overset{+1}{H}\overset{-1}{Br_{(aq)}} + \overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_{4(aq)}} \rightarrow \overset{+4}{S}\overset{-2}{O_{2(g)}} + \overset{0}{Br_{2(aq)}} + 2\ \overset{+1}{H_2}\overset{-2}{O_{(l)}}$

$1\ e^- / Br$ $2\ e^- / S$

$1\ e^- / HBr$ $2\ e^- / H_2SO_4$

(×2) (×1)

c. $2\ \overset{+7}{Mn}\overset{-2}{O_4}^-_{(aq)} + 6\ \overset{+1}{H^+}_{(aq)} + 5\ \overset{-2}{C}\overset{+1}{H_3}\overset{-2}{O}\overset{+1}{H_{(l)}} \rightarrow 2\ \overset{+2}{Mn^{2+}}_{(aq)} + 8\ \overset{+1}{H_2}\overset{-2}{O_{(l)}} + 5\ \overset{0}{C}\overset{+1}{H_2}\overset{-2}{O_{(aq)}}$

$5\ e^- / Mn$ $2\ e^- / C$

$5\ e^- / MnO_4^-$ $2\ e^- / CH_3OH$

(×2) (×5)

d. $2\ \overset{+5}{I}\overset{-2}{O_3}^-_{(aq)} + 5\ \overset{+1}{H}\overset{+4}{S}\overset{-2}{O_3}^-_{(aq)} \rightarrow 5\ \overset{+6}{S}\overset{-2}{O_4}^{2-}_{(aq)} + 3\ \overset{+1}{H^+}_{(aq)} + \overset{0}{I_{2(aq)}} + \overset{+1}{H_2}\overset{-2}{O_{(l)}}$

$5\ e^- / I$ $2\ e^- / S$

$5\ e^- / IO_3^-$ $2\ e^- / HSO_3^-$

(×2) (×5)

e. $\overset{0}{I_{2(aq)}} + \overset{+1\,+4\,-2}{HSO_3^-(aq)} + \overset{+1\,-2}{H_2O_{(l)}} \rightarrow \mathbf{2}\,\overset{-1}{I^-(aq)} + \overset{+6\,-2}{SO_4^{2-}(aq)} + \mathbf{3}\,\overset{+1}{H^+(aq)}$

$1\,e^-/I$ $2\,e^-/S$

$2\,e^-/I_2$ $2\,e^-/HSO_3^-$

f. $\mathbf{4}\,\overset{-3\,+1}{NH_3(g)} + \mathbf{7}\,\overset{0}{O_{2(g)}} \rightarrow \mathbf{4}\,\overset{+4\,-2}{NO_{2(g)}} + \mathbf{6}\,\overset{+1\,-2}{H_2O_{(g)}}$

$7\,e^-/N$ $2\,e^-/O$

$7\,e^-/NH_3$ $4\,e^-/O_2$

$(\times 4)$ $(\times 7)$

g. $\overset{-2\,+1\,-2\,+1}{C_2H_5OH_{(l)}} + \mathbf{4}\,\overset{+1}{H^+(aq)} + \mathbf{4}\,\overset{+5\,-2}{NO_3^-(aq)} \rightarrow \overset{0\,+1\;0\,-2\,-2\,+1}{CH_3COOH(aq)} + \mathbf{4}\,\overset{+4\,-2}{NO_{2(g)}} + \mathbf{3}\,\overset{+1\,-2}{H_2O_{(l)}}$

$2\,e^-/C$ $1\,e^-/N$

$4\,e^-/C_2H_5OH$ $1\,e^-/NO_3^-$

$(\times 1)$ $(\times 4)$

h. This is a case where the **same** element in the reactants (Cl) has some atoms which are oxidized and some atoms which are reduced – this is called auto-oxidation. To handle this kind of situation, write two of the starting reactants and then work from there.

$\mathbf{2}\,\overset{+1\,-2}{ClO^-(aq)} + \overset{+1\,-2}{ClO^-(aq)} \rightarrow \overset{+5\,-2}{ClO_3^-(aq)} + \mathbf{2}\,\overset{-1}{Cl^-(aq)}$

loss of $4\,e^-$

gain of $2\,e^-$

$2\,e^-/Cl$ $4\,e^-/Cl$

$2\,e^-/ClO^-$ $4\,e^-/ClO^-$

$(\times 2)$ $(\times 1)$

Overall, then, you have the following equation:

$$\mathbf{3}\,ClO^-(aq) \rightarrow ClO_3^-(aq) + \mathbf{2}\,Cl^-(aq)$$

5. **Textbook question 46:**

a. $1 \times 2 \text{ mol } e^- = 2 \text{ mol } e^-$ transferred in total

b. $1 \text{ mol } e^- \times 2 = 2 \text{ mol } e^-$ transferred
or $2 \text{ mol } e^- \times 1 = 2 \text{ mol } e^-$ transferred

c. $5 \text{ mol } e^- \times 2 = 10 \text{ mol } e^-$ transferred
or $2 \text{ mol } e^- \times 5 = 10 \text{ mol } e^-$ transferred

d. $5 \text{ mol } e^- \times 2 = 10 \text{ mol } e^-$ transferred
or $2 \text{ mol } e^- \times 5 = 10 \text{ mol } e^-$ transferred

e. $2 \text{ mol } e^- \times 1 = 2 \text{ mol } e^-$ transferred

f. $7 \text{ mol } e^- \times 4 = 28 \text{ mol } e^-$ transferred
or $4 \text{ mol } e^- \times 7 = 28 \text{ mol } e^-$ transferred

g. $4 \text{ mol } e^- \times 1 = 4 \text{ mol } e^-$ transferred
or $1 \text{ mol } e^- \times 4 = 4 \text{ mol } e^-$ transferred

h. $2 \text{ mol } e^- \times 2 = 4 \text{ mol } e^-$ transferred
or $4 \text{ mol } e^- \times 1 = 4 \text{ mol } e^-$ transferred

6. $2\,\overset{+6}{\underset{+6}{Cr_2}}\overset{-2}{O_7}^{2-}_{(aq)} + 16\,\overset{+1}{H^+}_{(aq)} + 3\,\overset{-2}{\underset{-2}{C_2}}\overset{+1}{H_5}\overset{-2}{O}\overset{+1}{H}_{(aq)} \rightarrow 4\,\overset{+3}{\underset{+3}{Cr^{3+}}}_{(aq)} + 11\,\overset{+1}{H_2}\overset{-2}{O}_{(l)} + 3\,\overset{0}{\underset{0}{C}}\overset{+1}{H_3}\overset{0}{C}\overset{-2}{O}\overset{-2}{O}\overset{+1}{H}_{(aq)}$

$3\ e^- / Cr$ $\qquad$ $2\ e^- / C$

$6\ e^- / Cr_2O_7^{2-}$ $\qquad$ $\boxed{4\ e^- / C_2H_5OH}$ ← C_2H_5OH is the RA. You need this information.

(×2) $\qquad$ (×3)

$$\text{mol } e^- = 35.2 \cancel{\text{ g } C_2H_5OH} \times \frac{1 \cancel{\text{ mol } C_2H_5OH}}{46.08 \cancel{\text{ g } C_2H_5OH}} \times \frac{4 \text{ mol } e^-}{1 \cancel{\text{ mol } C_2H_5OH}}$$

$$= 3.06 \text{ mol } e^-$$

When 35.2 g of reducing agent reacts, 3.06 mol of electrons are transferred.

7. a.

$$10\ e^- + 12\ H^+_{(aq)} + 2\ ClO_3^-{}_{(aq)} \rightarrow Cl_{2(g)} + 6\ H_2O_{(l)}$$

$$2\left(2\ H_2O_{(l)} + Cl^-_{(aq)} \rightarrow ClO_{2(aq)} + 4\ H^+_{(aq)} + 5\ e^-\right)$$

$$4\ H^+_{(aq)} + 2\ ClO_3^-{}_{(aq)} + 2\ Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2\ H_2O_{(l)} + 2\ ClO_{2(aq)}$$

or

$$2\left(1\ e^- + 2\ H^+_{(aq)} + ClO_3^-{}_{(aq)} \rightarrow ClO_{2(aq)} + H_2O_{(l)}\right)$$

$$2\ Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2\ e^-$$

$$4\ H^+_{(aq)} + 2\ ClO_3^-{}_{(aq)} + 2\ Cl^-_{(aq)} \rightarrow 2\ ClO_{2(aq)} + 2\ H_2O_{(l)} + Cl_{2(g)}$$

b.

$$3\left(Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2\ e^-\right)$$

$$2\left(3\ e^- + 3\ H^+_{(aq)} + HNO_{3(aq)} \rightarrow NO_{(g)} + 2\ H_2O_{(l)}\right)$$

$$3\ Cu_{(s)} + 6\ H^+_{(aq)} + 2\ HNO_{3(aq)} \rightarrow 3\ Cu^{2+}_{(aq)} + 2\ NO_{(g)} + 4\ H_2O_{(l)}$$

c.

$$5\left(Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + 1\,e^-\right)$$

$$5\,e^- + 8\,H^+_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}$$

$$5\,Fe^{2+}_{(aq)} + 8\,H^+_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow 5\,Fe^{3+}_{(aq)} + Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}$$

Notes

If you look at the half-reactions you generated in this question and compare them to the ones on the redox table, you will see they are the same. This tells you that you can use the redox table or generate your own half-reactions to balance redox reactions, whichever is most convenient.

8. The legal limit in Canada for blood alcohol content while driving is 0.08 mg/100 mL.

9. Breathalysers are used as a screening device to detect suspected impaired drivers, since it is very difficult to obtain blood samples.

10. A breathalyser works as follows:

- A breath sample is taken into the breathalyser and any alcohol present reacts with a known excess quantity of acidified potassium dichromate.
- The unreacted potassium dichromate is detected (by its orange colour) by the breathalyser and is converted to a blood alcohol content readout in mg/100 mL.

11.

$$Cr_2O_7^{2-}{}_{(aq)} + C_2H_5OH_{(aq)} \rightarrow Cr^{3+}_{(aq)} + CH_3COOH_{(aq)} \text{ (unbalanced)}$$

$$2\left(6\,e^- + 14\,H^+_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} \rightarrow 2\,Cr^{3+}_{(aq)} + 7\,H_2O_{(l)}\right)$$

$$3\left(H_2O_{(l)} + C_2H_5OH_{(aq)} \rightarrow CH_3COOH_{(aq)} + 4\,H^+_{(aq)} + 4\,e^-\right)$$

$$16\,H^+_{(aq)} + 2\,Cr_2O_7^{2-}{}_{(aq)} + 3\,C_2H_5OH_{(aq)} \rightarrow 4\,Cr^{3+}_{(aq)} + 11\,H_2O_{(l)} + 3\,CH_3COOH_{(aq)}$$

Note: You have seen this reaction balanced using a different method on page 378 of *Nelson Chemistry*.

12.

$$16\,H^+_{(aq)} + 2\,Cr_2O_7^{2-}{}_{(aq)} + 3\,C_2H_5OH_{(aq)} \rightarrow 4\,Cr^{3+}_{(aq)} + 11\,H_2O_{(l)} + 3\,CH_3COOH_{(aq)}$$

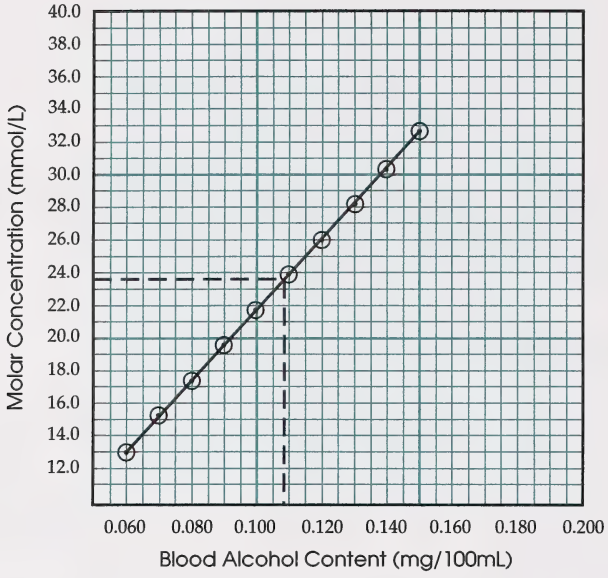
3.66 mL	0.500 mL
2.15 mmol/L	C = ?

$$C_{(C_2H_5OH)} = 3.66\ \cancel{mL\ Cr_2O_7^{2-}} \times \frac{2.15\ \cancel{mmol\ Cr_2O_7^{2-}}}{1\ \cancel{L\ Cr_2O_7^{2-}}} \times \frac{3\ mmol\ C_2H_5OH}{2\ \cancel{mmol\ Cr_2O_7^{2-}}} \times \frac{1}{0.500\ \cancel{mL}\ C_2H_5OH}$$

$$= 23.6\ mmol/L$$

The concentration of the ethanol was 23.6 mmol/L.

13. a. **Molar Concentration of Ethanol Versus Blood Alcohol Content**



From the graph, the molar concentration of 23.6 mmol/L of ethanol gives a blood alcohol content of 0.108 mg/100 mL.

b. Yes, the driver was legally impaired because the blood alcohol content exceeded the legal limit of 0.08 mg/100 mL.

Section 2: Follow-up Activities

Extra Help

1. a. ii
   b. iii
   c. i
   d. iv
   e. v

2. The $MnO_4^-{}_{(aq)}$ ion is a deep purple or violet. When it is reduced to the $Mn^{2+}{}_{(aq)}$ ion its colour would change to pink.

3. Another oxidizing agent other than the $MnO_4^-{}_{(aq)}$ ion which undergoes a colour change from orange to green when it is reduced would be $\mathbf{Cr_2O_7^{2-}{}_{(aq)}}$.

4. Trial 1 = 9.7 mL − 0.5 mL = 9.2 mL

Trial 2 = 18.5 mL − 9.7 mL = 8.8 mL

Trial 3 = 27.6 mL − 18.5 mL = 9.1 mL

$$\text{Average} = \frac{(9.2 + 8.8 + 9.1)\ \text{mL}}{3}$$
$$= 9.0\ \text{mL}\ (9.033\ \text{mL})$$

The average volume of acidified $KMnO_{4(aq)}$ solution was 9.0 mL.

5. OA SOA OA OA OA

$K^+_{(aq)}$, $MnO_4^-{}_{(aq)}$, $H^+_{(aq)}$, $Sn^{2+}_{(aq)}$, $Cl^-_{(aq)}$, $H_2O_{(l)}$

SRA RA RA RA

Note: Since there is a 1:1 ratio between $KMnO_{4(aq)}$ and $MnO_4^-{}_{(aq)}$, $C_{(KMnO_4)} = C_{(MnO_4^-)}$

$$2\left(MnO_4^-{}_{(aq)} + 8\,H^+_{(aq)} + 5\,e^- \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}\right)$$
$$5\left(Sn^{2+}_{(aq)} \rightarrow Sn^{4+}_{(aq)} + 2\,e^-\right)$$

$$2\,MnO_4^-{}_{(aq)} + 16\,H^+_{(aq)} + 5\,Sn^{2+}_{(aq)} \rightarrow 2\,Mn^{2+}_{(aq)} + 5\,Sn^{4+}_{(aq)} + 8\,H_2O_{(l)}$$

$v = 9.033$ mL $\quad C = 0.555$ mol/L

$C = ?$ $\quad v = 10.0$ mL

$$C_{(MnO_4^-)} = 10.0\ \cancel{\text{mL Sn}^{2+}} \times \frac{0.555\ \cancel{\text{mol Sn}^{2+}}}{1\ \cancel{\text{L Sn}^{2+}}} \times \frac{2\ \text{mol}\ \cancel{MnO_4}}{5\ \cancel{\text{mol Sn}^{2+}}} \times \frac{1}{9.033\ \cancel{\text{mL}}\ \cancel{MnO_4}}$$
$$= 0.25\ \text{mol/L}$$

The concentration of $MnO_4^-{}_{(aq)}$, and therefore $KMnO_{4(aq)}$, was 0.25 mol/L.

6. The list of species is as follows:

$$\overset{-2}{S^{2-}} \text{ or } H_2\overset{-2}{S},\ \overset{0}{S_8},\ \overset{+4}{S}O_2,\ H_2\overset{+6}{S}O_4$$

7. a. The following equation will be balanced by devising half-reactions. You may choose to balance the equation by assigning oxidation numbers. The end result should be the same.

$$Cr^{3+}_{(aq)} + BiO_3^-{}_{(aq)} \rightarrow Cr_2O_7^{2-}{}_{(aq)} + Bi^{3+}_{(aq)}\ \text{(unbalanced)}$$

$$7\,H_2O_{(l)} + 2\,Cr^{3+}_{(aq)} \rightarrow Cr_2O_7^{2-}{}_{(aq)} + 14\,H^+_{(aq)} + 6\,e^-$$
$$3\left(2\,e^- + 6\,H^+_{(aq)} + BiO_3^-{}_{(aq)} \rightarrow Bi^{3+}_{(aq)} + 3\,H_2O_{(l)}\right)$$

$$2\,Cr^{3+}_{(aq)} + 4\,H^+_{(aq)} + 3\,BiO_3^-{}_{(aq)} \rightarrow Cr_2O_7^{2-}{}_{(aq)} + 3\,Bi^{3+}_{(aq)} + 2\,H_2O_{(l)}$$

b. $4\ \overset{0}{Zn}_{(s)} + 4\ \overset{+1}{H_2}\ \overset{+6}{S}\ \overset{-2}{O_4}_{(aq)} + \overset{+1}{H_2}\ \overset{+6}{S}\ \overset{-2}{O_4}_{(aq)} \rightarrow 4\ \overset{+2}{Zn}\ \overset{+6}{S}\ \overset{-2}{O_4}_{(aq)} + \overset{+1}{H_2}\ \overset{-2}{S}_{(aq)} + 4\ \overset{+1}{H_2}\ \overset{-2}{O}_{(l)}$

gain $8\ e^-$

lose $2\ e^-$

$2\ e^- / Zn$ (×4)

$8\ e^- / S$

$8\ e^- / H_2SO_4$ (×1)

$4\ Zn_{(s)} + 5\ H_2SO_{4(aq)} \rightarrow 4\ ZnSO_{4(aq)} + H_2S_{(aq)} + 4\ H_2O_{(l)}$

Note: Two separate H_2SO_4's are written here because some sulphur atoms change in oxidation numbers, while others remain unchanged. The total H_2SO_4 count is shown in the bottom equation.

8. a. $2\ Sr_{(s)} + O_{2(g)} \rightarrow 2\ SrO_{(s)}$

$Sr_{(s)} \rightarrow Sr^{2+}_{(aq)} + 2\ e^-$

RA

$4\ e^- + O_{2(g)} \rightarrow 2\ O^{2-}_{(aq)}$

OA

b. $2\ Na_{(s)} + 2\ H_2O_{(l)} \rightarrow H_{2(g)} + 2\ NaOH_{(aq)}$

$Na_{(s)} \rightarrow Na^+_{(aq)} + 1\ e^-$

RA

$2\ e^- + 2\ H_2O_{(l)} \rightarrow H_{2(g)} + 2\ OH^-_{(aq)}$

OA

c. $Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2\ e^-$

RA

$2\ H^+_{(aq)} + 2\ e^- \rightarrow H_{2(g)}$

OA

Enrichment

1. $\overset{5}{H}\overset{}{N}O_3 + \overset{0}{Zn}_{(s)} \rightarrow ?$

OA **RA**

The oxidation number of nitrogen in nitric acid is +5. Since zinc is the reducing agent, the nitrogen must have to be the oxidizing agent. This means its oxidation number must decrease (be reduced).

Of the species given, the oxidation number of N in each of the following is lower than +5, and thus each of these could be produced:

- N = +2 in NO
- N = +4 in N_2O_4
- N = +4 in NO_2
- N = −3 in NH_4^+

Only N_2O_5 with an oxidation number for N equal to that in nitric acid (N = +5) could not be produced.

2. The analysis for Lab Exercise 12K is given on page 516 of your textbook *Nelson Chemistry*.

3. Reaction 1

$$Mn_{(s)} + HNO_{3(aq)} \rightarrow Mn^{2+}_{(aq)} + NO_{2(g)} \text{ (unbalanced)}$$

$$Mn_{(s)} \rightarrow Mn^{2+}_{(aq)} + 2\,e^-$$

$$2\left(1\,e^- + H^+_{(aq)} + HNO_{3(aq)} \rightarrow NO_{2(g)} + H_2O_{(l)}\right)$$

$$Mn_{(s)} + 2\,H^+_{(aq)} + 2\,HNO_{3(aq)} \rightarrow Mn^{2+}_{(aq)} + 2\,NO_{2(g)} + 2\,H_2O_{(l)}$$

Reaction 2

$$Mn^{2+}_{(aq)} + IO_4^-{}_{(aq)} \rightarrow MnO_4^-{}_{(aq)} + IO_3^-{}_{(aq)} \text{ (unbalanced)}$$

$$2\left(4\,H_2O_{(l)} + Mn^{2+}_{(aq)} \rightarrow MnO_4^-{}_{(aq)} + 8\,H^+_{(aq)} + 5\,e^-\right)$$

$$5\left(2\,e^- + 2\,H^+_{(aq)} + IO_4^-{}_{(aq)} \rightarrow IO_3^-{}_{(aq)} + H_2O_{(l)}\right)$$

$$3\,H_2O_{(l)} + 2\,Mn^{2+}_{(aq)} + 5\,IO_4^-{}_{(aq)} \rightarrow 2\,MnO_4^-{}_{(aq)} + 6\,H^+_{(aq)} + 5\,IO_3^-{}_{(aq)}$$

Reaction 3

$$MnO_4^-{}_{(aq)} + Fe^{2+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + Fe^{3+}_{(aq)} \text{ (unbalanced)}$$

$$5\,e^- + 8\,H^+_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)}$$

$$5\left(Fe^{2+}_{(aq)} \rightarrow Fe^{3+}_{(aq)} + 1\,e^-\right)$$

$$8\,H^+_{(aq)} + MnO_4^-{}_{(aq)} + 5\,Fe^{2+}_{(aq)} \rightarrow Mn^{2+}_{(aq)} + 4\,H_2O_{(l)} + 5\,Fe^{3+}_{(aq)}$$

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